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VIVIAN'S EVERYDAY CHEMISTRY
MOORE AND HALLIGAN'S PLANT PRODUCTION
TORMEY AND LAWRY'S ANIMAL HUSBANDRY

EDITED BY

KIRK LESTER HATCH, B.S.
PROFESSOR OF AGRICULTURAL EDUCATION
UNIVERSITY OF WISCONSIN, MADISON

AN ACHIEVEMENT OF APPLIED CHEMISTRY. The row of twelve apple trees on the left received an application of five pounds each of nitrate of soda and acid phosphate to the tree, and produced thirty barrels of apples. A similar row on the right, to which no chemicals were applied, produced only three barrels.

EVERYDAY CHEMISTRY

BY

ALFRED VIVIAN

DEAN OF THE COLLEGE OF AGRICULTURE OF
THE OHIO STATE UNIVERSITY

AMERICAN BOOK COMPANY
NEW YORK CINCINNATI CHICAGO
BOSTON ATLANTA

VIVIAN'S EVERYDAY CHEMISTRY

W. P. 3

GENERAL INTRODUCTION

THIS series of texts is based on the theory that the successful citizen should know the chemical, physical, and biological forces with which he has to contend; that he should understand the laws under which these forces operate; and that he should acquire some skill in directing them. He should ultimately become able to adjust and correlate these forces so as to bring them all under the orderly operation of economic law. In conformity with the above theory this series has been made to cover the following fundamental divisions:

1. *The science and art of chemistry as applied to everyday life, with special emphasis on household economics, soil fertility, and the relation of chemistry to plant and animal production.*
2. *The science and art of producing plants.*
3. *The production, care, and management of farm animals.*
4. *The proper balance and combination of these aspects of household economics and agricultural production, in the business management of the farm.*

KIRK LESTER HATCH.

PREFACE

The ordinary high-school course in chemistry in the past has consisted of a condensed treatment of the subject of inorganic chemistry, which was intended to prepare for the further study of that subject in college. The principal objection to a high-school course consisting entirely of inorganic chemistry lies in the fact that the ninety per cent of the pupils, who do not go to college and who pursue the subject no farther, are left with a very faint conception of the intimate relation which chemical phenomena bear to daily life. Such pupils are therefore likely to think of chemical changes as occurring only in beakers and in test tubes.

The reaction against the old type of high-school course in chemistry has resulted in the publication of a new type of textbooks for secondary schools. In these the authors have gone to the other extreme by attempting to present the applications of chemistry to daily life without any groundwork in the fundamental chemistry of the elements, a knowledge of which is necessary to the understanding of chemical compositions and reactions.

This text follows a middle course, and while the outstanding feature is its treatment of the applications of chemistry, inorganic and organic, the presentation is based on a brief study of the elements and their important compounds and reactions. The preparation of such a book is made difficult by the enormous mass of facts and theories from which must be selected those which are vital to the purpose of the text. Two questions have been asked regarding each fact and theory — (1) is this fact or theory essential to the understanding of any of the phenomena of daily life, or (2) is it necessary to explain some other fact or theory that is essential to such an understanding? Unless the answer to one of these questions is affirmative the subject in question is excluded from the text. The only exceptions to this rule are a few subjects proved by experience to be of especial value in holding the interest of the pupil. Much of the theory included in the older texts is excluded because of the feeling that a first course in chemistry should deal largely with facts and that the theoretical considerations should be left to a more advanced course. The chemistry of the elements has been confined to a score of elements that are of common occurrence. The study begins with the element itself if it is commonly known, but with a compound in case the element is not a familiar substance. In other words, the procedure is always from the known to the related unknown.

The small high school which cannot afford expensive equipment has been constantly in mind during the preparation of this text. The exercises call for no complicated apparatus; most of the experiments may be performed by means of simple homemade devices, or even in many cases by the use of kitchen utensils.

No one group of pupils should be required to use all the material in this book, since a wide range of topics has been treated in order to meet the varying needs of boys and girls. After the first thirty-five chapters have been completed the boys and girls may be separated and allowed to study those chapters which are of particular interest to each group. This book is especially suitable for use in the vocational courses in agriculture and home economics which are being introduced in many high schools.

Photographs for illustrations have been furnished by: Ohio Experiment Station, Frontispiece and Figs. 174, 225, 229, 231, 242, 244; Montana Experiment Station, Figs. 210, 211; Illinois Experiment Station, Fig. 236; L. H. Guldard, Figs. 92, 237; Jeffrey Manufacturing Company, Figs. 227, 228; Dr. E. V. McCullum, Fig. 159.

A WORD TO THE TEACHER

THE success of a course in chemistry depends largely upon the character of the laboratory work. Without the performance of laboratory exercises the study of chemistry degenerates into a mere act of memorizing. On the other hand, nothing is more pernicious than unsupervised laboratory work, since it not only wastes the time of the pupil but creates bad mental habits. The earnest teacher of chemistry will keep in close touch with the pupils throughout the course.

In small high schools where the class in chemistry does not include more than ten to fifteen pupils, it is a mistake to divide the time allotted to chemistry into recitation and laboratory periods. The teacher should be at liberty to use each period or part of a period for recitation or experimental work as the progress of the subject suggests; or better still, the experiment and the recitation should be coincident.

While it is desirable to have all the pupils perform each experiment, such a procedure is not absolutely essential to an understanding of the subject. Where the equipment is meager, very good results may be obtained by having the class gather around the laboratory table and having each student in turn perform an experiment while all members of the class take notes. This method gives the teacher an opportunity to have the class recite under the most favorable circumstances and especially to test the pupils' powers of inductive reasoning. It also tends to develop that kind of comradeship between pupil and teacher which is the successful teacher's greatest asset.

The teacher in the small high school needs no longer to hesitate to introduce chemistry because of the cost of apparatus. Most of the experiments in this text may be conducted with equipment which is at hand. By the exercise of a little ingenuity the remaining apparatus can be prepared from comparatively inexpensive materials. Common glass tumblers may be used in place of beakers when the application of heat is not necessary. Granite ware or porcelain-lined kitchen cups may be used in heating all liquids except very concentrated alkalies and acids. The only advantage in the use of glass beakers is that the experiment may be more readily observed. Much of the needed material may be collected locally by teacher and pupils, and many manufacturers are glad to furnish samples of their products gratuitously for use in the classroom.

It is not possible to eliminate from the course in chemistry the teacher, upon whom the success of the course depends more than upon any textbook or laboratory guide. For this reason the author has purposely omitted detailed descriptions of the laboratory experiments, leaving the specific directions to the teacher, who will have to give these directions in any event, no matter how complete the descriptions in the laboratory guide appear to be. The larger texts on chemistry, to which the teacher should refer, will give much assistance in arranging the details of the experiment.

The questions found in the exercises at the end of the chapters are merely suggestive, and many more will occur to the teacher. Their principal value is in enabling the pupils to test their own knowledge of the subject. The teacher should insist that the pupils look up every cross reference that is found in the text. Additional reading should be assigned on many of the topics, and it will be well to have each pupil report to the class upon certain topics which have been assigned as supplemental work. Most high schools own a good encyclopedia, and at least one article in the encyclopedia may be found which is germane to the matter discussed in each chapter of this text. The agricultural colleges of the several states, as well as the United States Department of Agriculture, publish bulletins for free distribution, many of which bear upon subjects discussed in this book. In many cases these bulletins appeal as much to the interest of people living in the city as to those in the country. Such pamphlets should be procured for the school library.

The high-school curriculum should fit the community in which the school is located, and the chemistry teacher has the opportunity of bringing his subject into close relationship with the daily life of the pupils. Chemical processes which are of unusual interest to the community should be emphasized, both in the classroom and by means of excursions to points where these processes may be observed. Chemistry should be made to assist the pupil in interpreting life. Above all, the teacher of chemistry should remember that the most important thing to be considered is the pupil and not the subject that is being taught.

INTRODUCTORY LABORATORY MANIPULATIONS

CHEMISTRY is essentially a laboratory study and is readily understood only when the statements of the text are illustrated by suitable exercises. As certain laboratory appliances and processes will be used repeatedly during the course, the student should first familiarize himself with them.

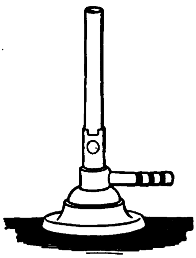

FIG. 1.

The Bunsen burner is used as a source of heat in most laboratories where gas is available. A common form of the burner is shown in Fig. 1. It is attached to the gas cock by a piece of rubber tubing. When the gas is turned on, the current of gas draws air through the holes at the bottom of the tube, and this mixture when lighted burns with an almost colorless flame, which is very hot and deposits no soot. The air supply may be reduced or entirely cut off by turning the ring at the bottom of the burner so that the holes in the tube are closed. As the air supply is lessened, the flame gradually becomes yellow and deposits soot. The colorless flame is used in all experiments unless the directions state otherwise.

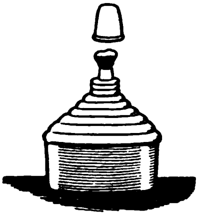

FIG. 2.

The alcohol lamp is ordinarily used in laboratories where a supply of gas is not available. The simple form shown in Fig. 2 is the most common. Such a lamp may be made in the laboratory from an ink bottle or an oil can and a little candle wicking. When a more intense heat is required the type of alcohol lamp shown in Fig. 3 or Fig. 4 may be used. In such a lamp the alcohol is converted into a vapor before it is burned. With proper precautions, the blow torch used by plumbers and painters (Fig. 5), in which gasoline vapor is burned, may be used when high temperatures are needed; but gasoline requires more careful handling than does alcohol.

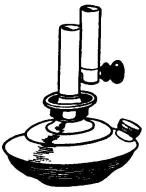
Fig. 3.

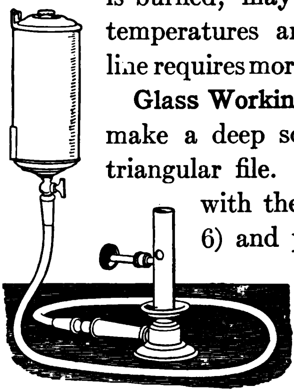
Fig. 4.

Glass Working. To cut a piece of glass tubing, make a deep scratch at the desired point with a triangular file. Grasp the tubing in both hands with the thumbs back of the scratch (Fig. 6) and pull the tubing apart, at the same time exerting a slight forward pressure of the thumbs. If the tubing does not break easily, make a deeper scratch with the file. Never use much pressure in breaking the tubing.

Smooth the ends of all glass tubing or glass rods by holding them in the flame of the burner until the glass fuses (Fig. 7). If desired, the end of the tubing may be closed entirely by holding it in the burner until the edges fuse together.

To bend glass tubing, place the wing top (Fig. 8) on the burner and hold the tubing in the upper part of the flame until the loose end begins to drop of its own weight. Then grasp the loose end

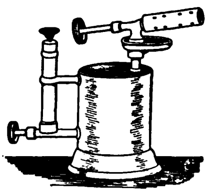
Fig. 5.

and bend to the desired angle. Avoid bends like *A* and *B* in Fig. 9, which are due to heating too small an area of the tube.

A glass tube may be drawn out by heating a small area in the point of the flame (Fig. 10) until the glass is quite soft and the walls of the tube are thickened. The tube must be rotated in the flame during heating. Remove the tube from the flame and draw it gently in a horizontal position until it is reduced to the desired size (Fig. 11).

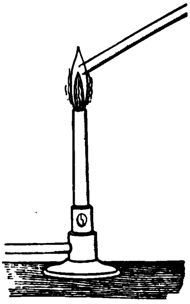

Fig. 7.

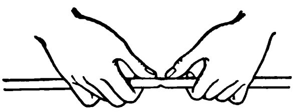

Fig. 6.

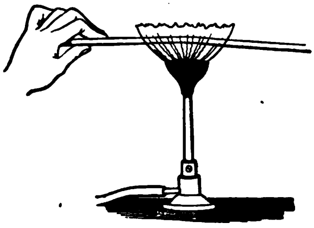

Fig. 8.

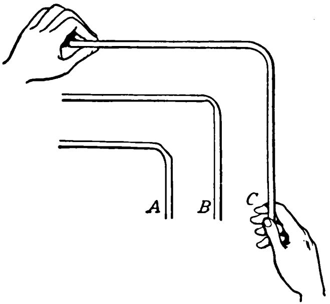

Fig. 9.

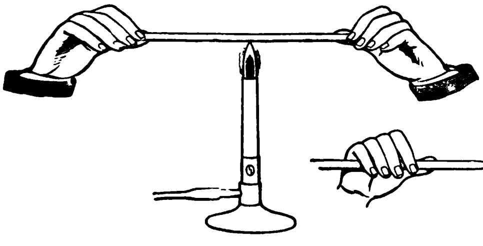

Fig. 10.

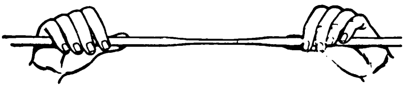

Fig. 11.

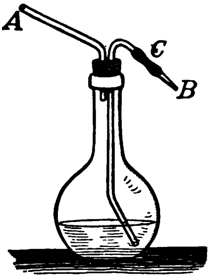

FIG. 12.

Wash Bottle. Make a wash bottle similar to Fig. 12. This will be found useful in many ways in the laboratory. By blowing in the tube (*A*) a small stream of water is forced out of the jet (*B*). A piece of rubber tubing at *C* permits the stream to be guided in any direction.

Holes in corks are usually made with a cork borer (Fig. 13). A round or rattail file may be used, the hole being filed out until it is the right size for the tubing. The tubing is more easily inserted if it is moistened. If the hole in the cork is too small, there is danger of breaking the tubing and seriously cutting the hands. Rubber corks with holes in them are now so easily obtained that they should be used when possible.

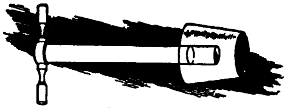

FIG. 13.

Heating Liquids in Test Tubes. Hold the test tube by means of a test tube holder or a band of paper as in Fig. 14. Heat with the point of the flame near the top of the liquid, but do not allow the flame to strike the glass above the liquid. Agitate slightly during heating. Be careful that the mouth of the test tube does not point toward any one in case the vapor forces the liquid out of the tube.

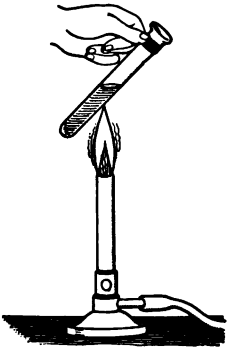

FIG. 14.

When heating liquids in glass beakers the same precaution must be observed against allowing the flame to strike the glass above the liquid. Beakers should be protected by placing them on a piece of wire gauze or asbestos board (Fig. 15). In some cases it is desirable to use a sand bath, which consists of a small pan containing sand (Fig. 16). This distributes the heat evenly and prevents breakage.

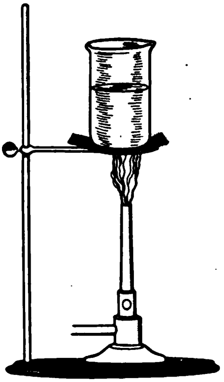

Fig. 15.

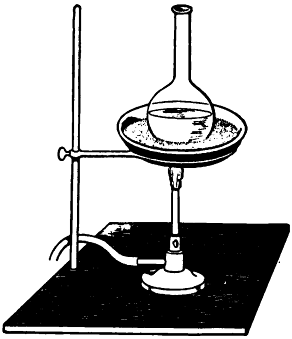

Fig. 16.

Evaporation is commonly performed in small porcelain evaporating dishes (Fig. 17). Wire gauze or the sand bath is frequently used with the evaporating dishes. If the substance is injured by high temperature, it is evaporated over a water bath (Fig. 18), in which case the evaporating dish is heated by the steam of the boiling water. The double

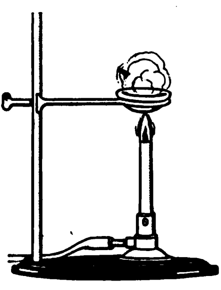

Fig. 17.

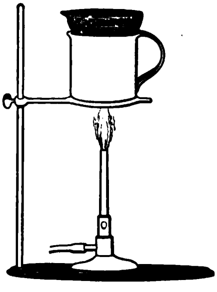

Fig. 18.

boiler is an example of the practical use of a water bath in the home.

Filtration is used to separate a solid from a liquid substance. It is usually performed by filtering the liquid through a specially prepared paper known as filter paper. This comes in circular pieces in various sizes, which are folded as shown in Fig. 19 so as to fit into a glass funnel (Fig. 20). When poured on the filter paper the liquid runs through, while the solid remains in the funnel.

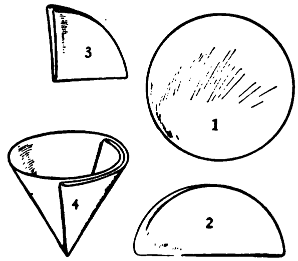

Fig. 19.

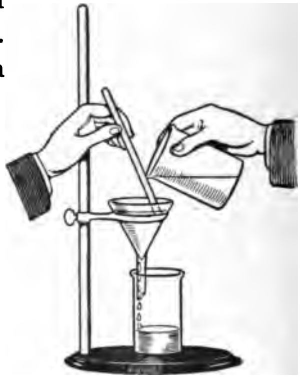

Fig. 20.

In transferring liquids from one vessel to another it is best to pour the liquid down a glass rod (Fig. 20) as this prevents danger of loss of the substance by splashing. Direct the stream against the side of the beaker or other vessel to which the liquid is being transferred. When pouring into a filter direct the stream against the side of the filter having the three layers of paper.

Fig. 21.

Keep the reagents pure by using extreme care to prevent contamination of any kind. Do not lay the stopper on the desk, but take it from the bottle as shown in Fig. 21, holding both the bottle and the stopper in the fingers as indicated in Fig. 22. Never pour any liquid back into the reagent bottle. A very little foreign substance in a reagent may spoil a future experiment. Use small quantities of all reagents. Many experiments are ruined by the use of too much material.

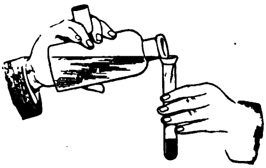
Fig. 22.

To collect gases the pneumatic trough is used (Fig. 23). Any deep pan will serve the purpose. The bottle used to collect the gas is filled with water and placed mouth downward in the pneumatic trough. The bottle is slightly tilted so that the gas may enter at *A* and displace the water in the bottle. Experiment in collecting gases by placing a piece of tubing under the bottle, as shown in the illustration, and gently blowing into it. To remove the bottle of gas from the trough slip a small piece of window glass beneath the mouth of the bottle (Fig. 24) while it is still under water. If the gas is heavier than air, the bottles should be stored mouth upward with the glass plate on top. If the gas is lighter than air, the bottles are kept mouth downward.

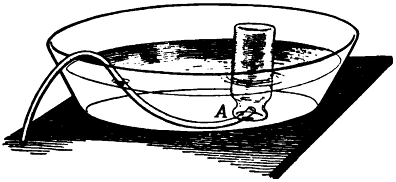

Fig. 23.

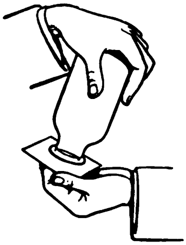
Fig. 24.

Some gases are so soluble in water that some other method than the one described above must be used in collecting

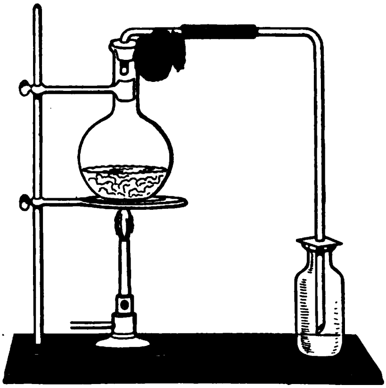

FIG. 25.

them. They may, in many cases, be collected over mercury, but the method more commonly used in the laboratory is by displacement of air. The gas is allowed to run into the collecting vessel until it has driven out all or practically all of the air. When the gas is heavier than air it is collected by downward displacement (Fig. 25). If lighter than air, the gas is collected by upward displacement (Fig. 26).

Cleanliness is absolutely necessary to successful laboratory work. All apparatus should be clean before use, and should be washed as soon as the experiment has been completed. Directions should be carefully followed in all the experiments, and no exercise should be started until the entire procedure is thoroughly understood by the student.

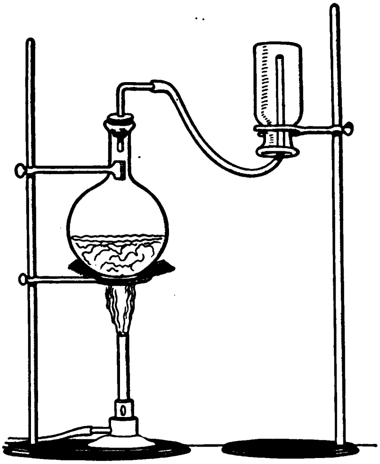

FIG. 26.

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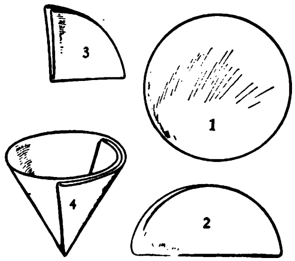

FIG. 19.

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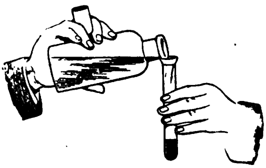


FIG. 22.

To collect gases the pneumatic trough is used (Fig. 23). Any deep pan will serve the purpose. The bottle used to collect the gas is filled with water and placed mouth downward in the pneumatic trough. The bottle is slightly tilted so that the gas may enter at *A* and displace the water in the bottle. Experiment in collecting gases by placing a piece of tubing under the bottle, as shown in the illustration, and gently blowing into it. To remove the bottle of gas from the trough slip a small piece of window glass beneath the mouth of the bottle (Fig. 24) while it is still under water. If the gas is heavier than air, the bottles should be stored mouth upward with the glass plate on top. If the gas is lighter than air, the bottles are kept mouth downward.

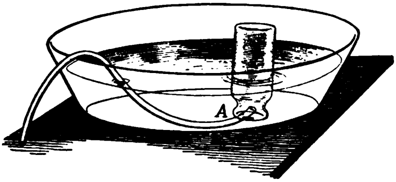


FIG. 23.

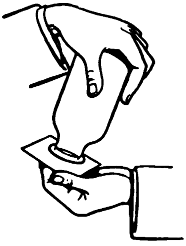


FIG. 24.

Some gases are so soluble in water that some other method than the one described above must be used in collecting easily be shown by evaporating a small quantity to dryness in a porcelain evaporating dish, as illustrated in Fig. 27. If like amounts of water from various sources are evaporated, it will be found that there is a great variation in the amount of solid substance left in the evaporating dishes. The nature of the materials obtained from well and spring water depends, evidently, upon the character of the rock and soil in which the well or spring is located. Granite rocks are very insoluble, while limestone is much more readily dissolved. Water issuing from the former contains very little dissolved substance, whereas it is a well-known fact that the water in limestone regions is heavily charged with mineral matter.

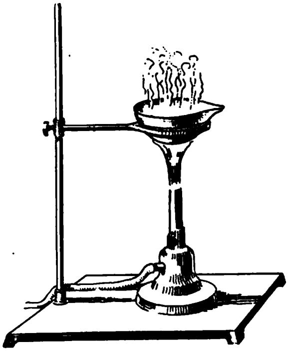


FIG. 27. — **Evaporating water to show the presence of dissolved mineral matter.**

3. Rain Water the Purest Natural Water. The purest natural water is rain water, but even that is not perfectly pure. The atmosphere always contains more or less dust and smoke, and when rain falls it carries these substances down with it. Rain water in the open country is purer than that in or near the cities, and if the water is not collected until the air has been washed by the rain for some minutes, comparatively pure water may be obtained.

4. Distillation of Water. Pure water is obtained by boiling well water or hydrant water and condensing the steam. This process is known as *distillation* and may conveniently be carried out in the apparatus shown in Fig. 28. The water is boiled in the flask *A*, and the steam passing into the inner tube of the condenser *B* is cooled and changed back to water, which is collected in the receiving flask *C*. A current of cold water is kept running through the condenser, entering at *D* and flowing out at *E*. This water cools the inner tube of the condenser sufficiently to cause the condensation of the steam entering from the flask *A*. The water which collects in *C* is known as *distilled water* and is quite pure, since the impurities of the original water remain in the boiling flask. The first fifty cubic centimeters which pass over should be rejected, as gases and other volatile substances may distill over with the first portions of water, and the inside of the condenser may not be entirely free from soluble materials. If proper precautions have been observed, the distilled water made in this way will leave no residue when evaporated in a porcelain dish.

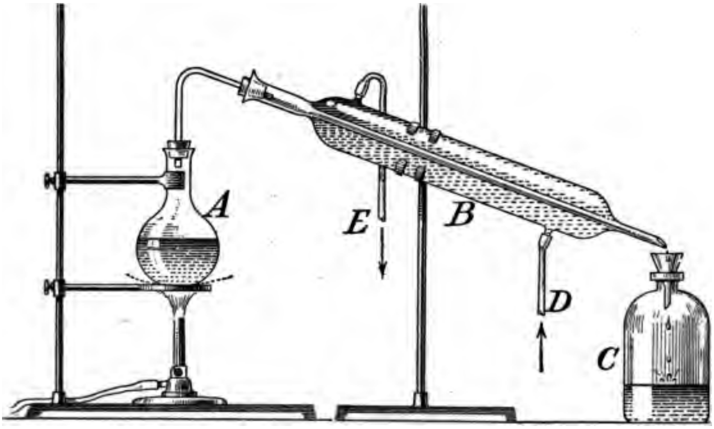


FIG. 28. — Apparatus for the distillation of water.

5. Properties of Water. Pure water is tasteless and odorless. In small quantities it seems to be colorless, but when viewed in deep layers it becomes apparent that it has a blue color. This is illustrated in the beautiful blue color of many mountain lakes, which consist of almost pure water derived from the melting snow.

6. Three States of Water. When water is boiled it is gradually changed into a colorless and invisible vapor, which, upon cooling, is again converted into liquid water. If the vapor escapes into the cooler air, it is partially condensed and forms what is popularly known as steam. That true steam, or water vapor, is invisible can be shown by boiling water in a flask. It will be found that nothing can be seen in the upper part of the flask although it must be full of steam. It is only as the steam escapes into the air and is condensed to small droplets of water that it can be seen.

Most substances contract when they are cooled. Water when cooled follows the general rule to four degrees centigrade, whereupon it begins to expand. At 4° C. water reaches its *maximum density*. At the moment the water freezes, a considerable increase in volume takes place, and the resulting ice has a density not much more than nine tenths that of the water from which it was formed. This explains why ice is always found on top of the water. If the volume contracted as freezing took place, ice would sink to the bottom, and the lakes and rivers would be frozen to a solid mass of ice. This expansion of the water upon freezing also accounts for the fact that ice is always pushed up on the banks of the lakes and rivers. The great force exerted by the expansion of water at the moment of freezing is well known, and nearly every one can recall an experience

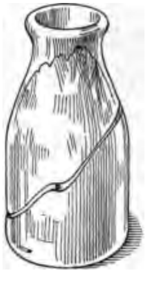


FIG. 29.—**Bottle broken by the freezing of the contained water.**

with the bursting of a water pipe, a bottle (Fig. 29), a bucket or of some other vessel, due to the freezing of the contained water. This force is also an important factor in the weathering of rocks and in the formation of soils (483).

7. Water used to Establish Standards. Pure water is used to establish many of the scientific standards of measurement. The two fixed points of the thermometer are the boiling and the freezing points of water. To graduate a thermometer the bulb is placed in melting ice, and the height of the mercury is marked. This point is the *zero* of the centigrade scale, or *thirty-two degrees* on the Fahrenheit scale (Fig. 30). The thermometer is then immersed in boiling water and the height of the mercury again marked. This point is marked 100 degrees centigrade, or 212 degrees Fahrenheit. The space between these marks is then divided into one hundred equal spaces for the centigrade scale and into 180 for the Fahrenheit. Since the boiling and the freezing points of liquids vary with the atmospheric pressure, the graduation described above must be made at sea level, or be corrected for the difference in pressure. Increased pressure raises the boiling point and lowers the freezing point, while decreased pressure has the opposite effect. That water under decreased pressure boils at a lower temperature can readily be shown by boiling water in a flask closed with a two-hole rubber cork, in one hole of which is placed a thermometer, and in the other a tube that is connected with an air pump (Fig. 31). If the air is

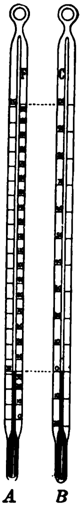


FIG. 30. — Fahrenheit (*A*) and centigrade (*B*) thermometers showing the relation of the scales.

partially exhausted from the flask, the water will be found to boil much below 100° C. The same result may be obtained by heating water in a flask, and, while the water is boiling hard, closing the flask with a rubber stopper, and removing the flask from the flame. If now a stream of water is run over the upper part of the flask (Fig. 32), the steam will be condensed, a partial vacuum will be produced, and the water will again boil. This may be continued until the temperature of the water is much below its normal boiling point. On the top of Pikes Peak water boils at such a low temperature that an egg cannot be cooked hard in it.

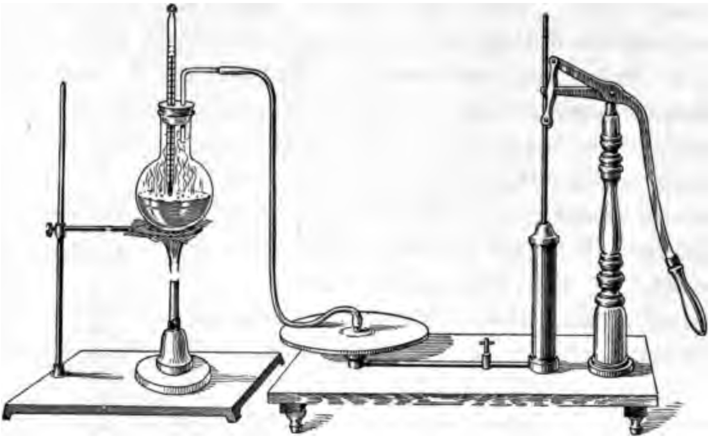


FIG. 31. — **Apparatus to show effect of reduced pressure on boiling point of water.**

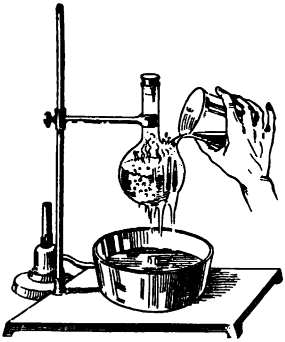


FIG. 32. — **A simple way of showing the effect of reduced pressures on the boiling point.**

The gram, the standard of weight of the metric system, is the weight of one cubic centimeter of water at its maximum density; that is, at 4° C. Water is also used as a standard of density; for the relative density, or *specific gravity,* of any substance is determined by dividing its weight by that of an equal volume of water. The standard for the measurement of a quantity of heat is the *calorie,* which is the amount of heat required to raise the temperature of one gram of water 1° C. This is sometimes called the small calorie. The large calorie is the amount of heat required to raise 1000 grams of water 1° C.

8. Heat of Fusion and Vaporization. When ice is heated it melts, but if a thermometer is placed in the melting ice (Fig. 33), and the mass is kept thoroughly stirred, it will be found that the temperature of the mass does not rise until all the ice has been melted. Although much heat has been applied, it has all been used to melt the ice and not to raise the temperature. The heat required merely to melt a substance is termed *heat of fusion.* When the water freezes, the same amount of heat is given off. Similarly a large amount of heat is required to change water into vapor, or steam, and this is known as the *heat of vaporization* of the water. Whenever water evaporates it absorbs an amount of heat equal to its heat of vaporization and, consequently, cools both the air in the vicinity and the surface from which evaporation takes place. This explains why sprinkling the floor on a hot day cools the room. In India, wet cloths are

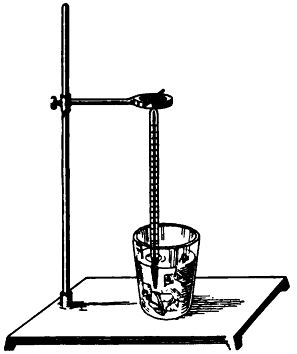


FIG. 33. — The temperature of the water does not rise until all the ice is melted.

frequently hung in doorways, so that the air entering the house will be cooled by evaporation of the water. Drinking water in hot climates is often stored in semi-porous earthenware jars (Fig. 34) in order that the water oozing through may evaporate from the surface and thus cool the jar and its contents.

FIG. 34. — Semi-porous water jars on sale in a bazaar in Allahabad, India.

9. Water a Poor Conductor of Heat. It may be shown that water is a poor conductor of heat by applying heat to the top of a test tube full of water (Fig. 35). The water in the upper portion of the tube can be made to boil while the bottom is still cool. To raise the temperature of a large body of water it is necessary to apply heat at the bottom. When the water next to the fire becomes heated, it expands, becomes less dense, and rises to the top (Fig. 36). The cooler and heavier liquid streams down to replace it, and thus a system of currents is set up that gradually distributes the heat throughout the whole mass. This statement applies to other liquids and to gases as well as to water.

An interesting series of changes takes place when a pond or other body of water freezes over. As the air above the water becomes colder, heat is given off from the surface of the water, and the cold water from above streams down to the bottom of the pond and forces the warmer water to come to the surface. This continues until all the water in the pond is cooled to 4° C. Upon further cooling the water at the surface expands, thus becoming lighter, and no longer moves downward. When the surface is cooled to 0° C., the water freezes, giving off an amount of heat equal to its heat of fusion, and the temperature of the water immediately below the ice is temporarily raised. The layer of ice increases in thickness due to the loss of heat from the surface until it is sufficiently thick to prevent further radiation of heat from the water beneath, ice being a relatively poor conductor of heat. It thus happens in deep ponds that the water at the bottom does not fall to the freezing point, even in very cold weather.

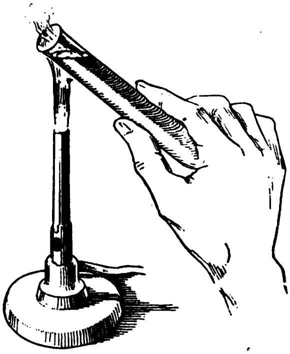


FIG. 35. — **Showing that water is a poor conductor of heat.**

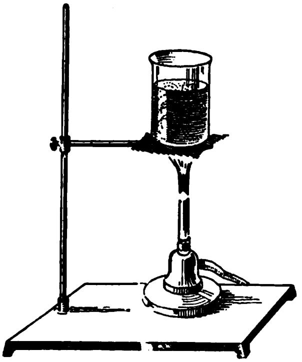


FIG. 36. — **Currents of warm water move upward as the bottom layers are heated.**

10. Water Has High Specific Heat. Water requires more heat to make it hot than does any other substance. If equal amounts of heat are applied to equal masses of water and mercury, for instance, the mercury gets hot much more rapidly than the water. Water, therefore, is said to have a greater capacity for heat, or a greater *specific heat,* than other substances. This property of water is of importance, for substances that heat slowly also cool slowly. This is one reason why hot water and steam are used to heat houses, and explains why large bodies of water have a moderating effect upon the climate of their vicinity.

EXERCISES

Ex. 1. Evaporate 10 cc. each of well water and rain water to dryness by heating in glass dishes or watch glasses on a sand bath (Fig. 37). What remains in the dish? Which water gave the larger residue? Which sample of water was the purer? Was either one perfectly pure? How did the residue in the dish get into the water? From which well should you expect the most residue, one located in granite, sandstone, or limestone? Why? Why does sea water contain large quantities of salt and other mineral matter? Where does the mineral matter in sea water come from? Explain the cycle of water in nature.

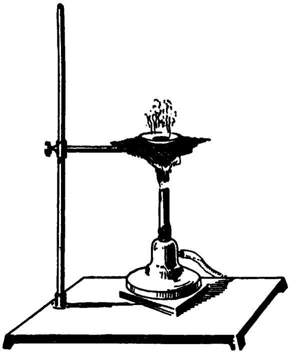


FIG. 37. — **Evaporating a small quantity of water on a watch glass.**

Ex. **2.** Arrange an apparatus according to Fig. 28. Half fill the flask *A* with the well water used in Ex. 1 and boil until 50 cc. of water is collected in *C*. Evaporate 10 cc. of the distilled water to dryness as above. Is there any residue? Compare with the well water before distillation. What has become of the mineral matter that was in the well water? How is pure water obtained?

(*Note.* If the condenser shown in Fig. 28 is not available, sufficient distilled water for this test can be produced with the apparatus shown in Fig. 38. Surround the test tube with the coldest water obtainable. If ice is available to cool this water, more distilled water can be produced.)

Ex. 3. Add a few crystals of copper sulphate to the water in flask *A* and distil a few cubic centimeters of water. What is the color of the water in the flask *A*? What is the color of the distilled water (sometimes called the *distillate*)? Explain the difference. In producing pure distilled water why should the first water that distilled over be rejected? State the properties of pure water.

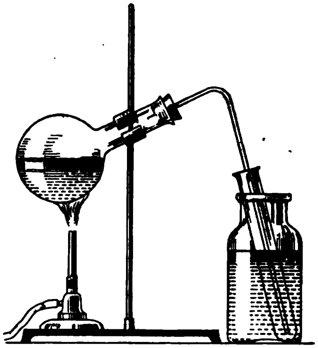


FIG. 38. — **A simple apparatus for producing a small quantity of distilled water.**

Ex. 4. Boil some water in a flask which is open at the top. Is the vapor in the upper part of the flask visible? Why does the "steam" become visible when the vapor passes into the air? Use a thermometer to determine the temperature of the water vapor (Fig. 39).

Ex. 5. If the weather is sufficiently cold, place a bottle full of water out of doors to freeze. Does water contract when cooled? What happens when the water freezes? Why does ice float on top of the water? What would happen to lakes if ice were heavier than water? At what temperature is water at its maximum density?

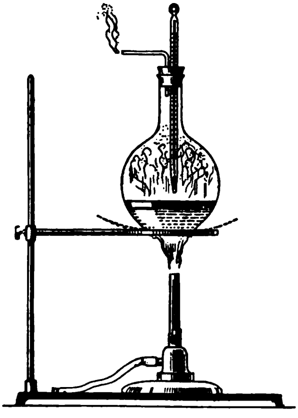


FIG. 39. — **Determining the temperature of water vapor.**

Ex. 6. Boil water in a partial vacuum created by an air pump, as suggested in the text. At what temperature does the water boil? At what temperature did it boil in Ex. 4? Explain the difference. What is the elevation of your school above sea level? Should you expect water to boil in your laboratory at 100° C. or below? Why?

Ex. 7. Fit a flask with a good rubber stopper and place in the stopper a short piece of glass tubing. Attach a piece of rubber tubing and a pinchcock to the upper end of the glass tubing. Boil the water in the flask until all air is driven out, and while the water is still boiling close the pinchcock and immediately remove the flame. Now run cold water over the upper part of the flask. Explain what happens to the water in the flask. Why does it take longer to cook vegetables by boiling at high altitudes than at sea level? If the flask used in this exercise were strong enough to risk boiling the water after the cork was inserted, would the boiling point of the water be raised or lowered? Why?

Ex. 8. Place a beaker or tin cup of water over a burner and watch the rise of temperature of the water by means of a thermometer. Does the temperature begin to rise immediately? Try the experiment again, using water with ice in it. Does the temperature rise immediately in this case? Explain the difference.

Ex. 9. Cool a pound of water (1 pint) to zero by immersing the vessel in a mixture of ice and salt until the desired temperature is reached; then remove from the cooling mixture. Have another pint of water heated to 80° and pour into the water at zero. Stir it quickly with the thermometer and read the temperature. What is the temperature of the mixture?

Weigh one pound of ice and pour over it a pint of water which has been warmed to 80°. Stir and read the temperature as soon as the ice has melted. What is the temperature in this case? How do you explain the difference in the two parts of this exercise? (Note that the ice used in this experiment must be at 0° C. Can ice be colder than zero? If all conditions are right for the above experiment, the mixture of water at 80° and water at zero will have a temperature of 40°, while the mixture of water at 80° and ice at zero will have a temperature of zero. Eighty times as much heat is required to melt a pound of ice as is required to warm a pound of water one degree.)

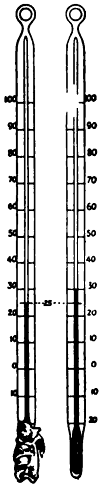


FIG. 40. — **Wet and dry bulb thermometers to show the effect of evaporation on the temperature.**

What is meant by heat of fusion? What becomes of the heat of fusion when the water freezes? When there is danger of freezing it is sometimes suggested that a tub of water be placed in the cellar. What is the theory of this recommendation?

Ex. 10. Compare two thermometers to see that they register alike. Wrap a piece of thin cloth around the bulb of one and moisten it with water (Fig. 40). Fan the two thermometers and observe the change in temperature registered. Is there any difference in the two thermometers? Explain. Why does sprinkling the floor lower the temperature of the room? Explain the use of porous jars to cool water in hot climates. Can you think of any other instance of the cooling effect of evaporating water?

Ex. 11. Fill a long test tube with water and heat the upper part over a burner until the water boils. Is the lower part of the tube hot? How do you explain it? Where should the heat be applied to warm water with the least fuel? Explain. Explain the changes that take place in the freezing over of a pond. What is meant by specific heat? How does the specific heat of water compare with that of other substances? Is there any connection between this fact and the use of hot water and steam in heating buildings? Explain the moderating effect which large bodies of water have upon climate.

CHAPTER II

WATER (*Continued*)

11. Water the Best Solvent. Water is the best-known solvent. In other words, it will dissolve more substances than any other liquid. There are some substances, however, which are insoluble in water. Substances vary greatly in solubility, ranging from those which will dissolve in a fraction of their own weight of water, to those whose solubility can scarcely be detected. Water dissolves other liquids and gases as well as solid substances. Liquids which are soluble in water are said to be *miscible* with water. Alcohol is miscible with water in all proportions, and so is sulphuric acid. Oils, on the other hand, are practically insoluble in water. The fact that air and other gases are dissolved in natural waters can be shown by gently heating the water in a glass vessel (Fig. 41), whereupon bubbles of gas will be seen to separate. The solubility of substances in water is affected by temperature. As a general rule, the solubility of solids increases with the rise of temperature, while that of gases decreases. Increase of pressure has a very marked effect upon the solubility of gases in water, for the amount

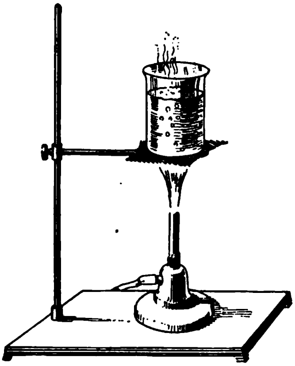


FIG. 41. — **Warming water to show the presence of dissolved gases.**

of gas dissolved is directly proportional to the pressure; that is, if the pressure upon the gas is doubled, twice as much will be dissolved by the water.

12. Dissolved Substances Raise the Boiling Point of Water. Substances dissolved in water raise its boiling point and lower its freezing point. Water in which salt is dissolved, for instance, has to be heated above 100° C. before it will boil, and it will not freeze until much below 0° C. This effect of the dissolved substance is made use of in a number of practical ways, some of which will be referred to later. When the water contains all of the substance it can dissolve, the solution is said to be *saturated.* From what has been said above it will be seen that a solution which is saturated at a low temperature will no longer be saturated if the temperature is raised, and that more of the substance may then be dissolved. Conversely, if the temperature of a saturated solution is lowered, some of the dissolved substance will be thrown out of the solution.

13. Water Accelerates Chemical Action. Water is also of interest to the chemist because it hastens chemical changes. In fact, most of the important chemical changes will not take place at all in the absence of water. If a small quantity of dry baking soda and dry powdered tartaric acid are mixed in a beaker, no change will take place. If, however, a little water is added, a marked change takes place, the most noticeable thing being the large amount of gas that is evolved. The rusting of iron is another example of a chemical change which occurs only in the presence of moisture.

14. Potable Waters. What is popularly meant by a pure water is one that is fit to use for drinking purposes. Such a water is technically called *a potable water.* Sea water with its large amount of salt is obviously unfit to drink, but it is seldom indeed that the mineral substances found in well or spring water are injurious. The chief source of danger in drinking water is that it may be contaminated (Fig. 42) with sewage, if the source of the water is near human habitations. The water from cesspools may enter the well or spring, or the city sewage may be emptied into the river. In either case disease-producing bacteria find their way into the drinking water. Typhoid fever and other diseases are often spread in this way. Wells should not be dug or drilled close to cesspools, for water from the surrounding soil is likely to drain into them. Deep wells are less likely to be contaminated than shallow ones, but in any case the upper part of the wells should be made water tight so that no surface water can enter. The water-tight casing should be carried down to a layer of clay or stone through which the surface water cannot penetrate. A spring at the foot of a hill is often contaminated from a cesspool on the hillside above.

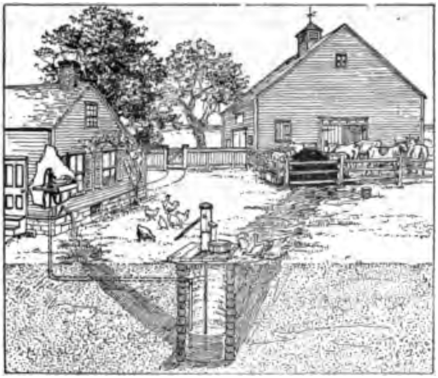


FIG. 42. — The contamination of water.

15. Boiled Water for Drinking. In case water is suspected of contamination the *only safe thing to do is to boil it, and thereby kill the bacteria.* Many filtering devices are offered for sale, but no small household filter can be depended upon to remove all the disease-producing bacteria.

In the case of city water supply large filtering beds may be constructed which will purify the water, but even these require the constant care of the expert engineer to keep them working effectively. Boiled water has a flat taste because of the fact that the air and other gases have been driven out by the boiling process. The taste of such water may be improved, however, by beating air into it after it has been thoroughly cooled. Rain water, or cistern water, collected in the open country, may be used for drinking purposes with safety, *if the rain is first allowed to wash the roof, and the cistern is carefully protected to prevent contamination.*

16. Hard Water. Water which contains a large amount of mineral matter is known as *hard water.* With such waters, soap, instead of forming suds, produces a curd which floats on top. Hardness is for the most part caused by lime compounds that are dissolved in the water. If the water is boiled, some of these lime compounds are made insoluble and separate, leaving the water less hard than it was. Hardness which can be removed by boiling is known as *temporary hardness,* while that which cannot be so removed is called *permanent hardness.* These terms will be more fully explained in a later chapter. Hard water is unsuited to household use because so much more soap is required with it than with soft water. This may be illustrated by a simple experiment. Equal quantities (25 cc.) of distilled water, rain water, fresh well water, and boiled well water are placed in four 8-ounce bottles. A one per cent solution of a pure soap (in alcohol) is added to each bottle, a drop at a time, and the bottle is vigorously shaken after each addition. It will be found that almost the first drop will make suds with the distilled water; that a little more will be required with the rain water; and that the well water will require much more than either the distilled or the rain water. The boiled well water will produce suds with less soap than the fresh well water.

Hard water is not satisfactory for use in boilers, for the mineral matter in such water forms a crust or scale on the inside of the boiler much like that found on the inside of a teakettle when well water is used. In limestone countries, especially, it becomes necessary to soften the water before it can be used in the boilers. This is done by the addition of boiler compounds made of various chemicals which will throw the lime or calcium salts out of solution. Borax, washing soda, quicklime, sodium phosphate, and other chemicals are used for this purpose.

17. Mineral Waters. Mineral waters are not necessarily heavily charged with mineral matter, but usually contain rather large amounts of one substance which gives them their peculiar characteristics. Some contain magnesium compounds, others are noted for the iron they contain, and still others are so charged with carbon dioxide that they effervesce like soda water. Sulphur water and lithia waters are also found in nature.

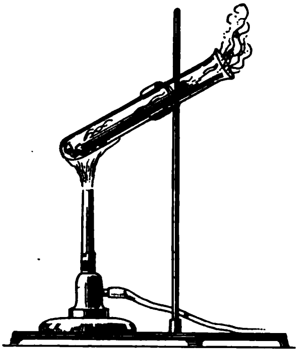


FIG. 43. — Showing the presence of water in organic matter.

18. Water in Organic Matter. The sources of water heretofore mentioned, such as wells, rivers, and lakes, are obvious; but water also exists under conditions not so apparent. It is well known that plants and animals use large quantities of water during their growth, and that some of the water remains in the organism. If a small quantity of grass, potato, turnip, corn, or a bit of lean meat is heated gently in a test tube, as shown in Fig. 43, the escaping water will be condensed in droplets at the top of the tube. Vegetables like beets, carrots, turnips, and potatoes are 90 per cent water. If the body of a calf weighing 150 pounds were completely dried, it would be found to weigh only about 50 pounds. In other words, such a body is nearly two thirds water.

19. Water of Crystallization. If a crystal of Glauber's salt (sodium sulphate) is heated in a test tube, water will be driven off and the crystal will change to a powder. It may be inferred from this experiment that the crystal contained water, and that the water was in some way necessary to produce the crystalline form. If a crystal of blue vitriol (copper sulphate) is heated, not only will water be driven off and the crystalline form be destroyed, but the blue color will disappear as well, leaving a grayish white mass. A little water added to this mass will restore the blue color, and if the material is dissolved in a small quantity of hot water and allowed to stand, blue crystals will again be formed. Water which thus forms a part of the crystal is known as *water of crystallization.* Not all crystals contain water, however, as some compounds crystallize without water of crystallization. Quartz and potassium dichromate are examples of such compounds.

20. Efflorescence and Deliquescence. If a bright crystal of washing soda is allowed to stand exposed to the air, it gradually gives off a part of its water of crystallization and crumbles to a white powder. Crystals which thus give off their water at ordinary temperatures are said to *effloresce.* Other substances have such a strong attraction for water that they will take it from the air until they actually dissolve in the absorbed water. Such materials (of which calcium chloride is a good example) are said to *deliquesce.* Substances which remove water so readily from the air will be found to be useful in drying gases.

21. Water May Be Decomposed. Water exists in three different forms — solid, liquid, and gaseous water. These differences, however, are purely physical. Ice may be melted to water, and the water heated to steam and distilled, and, under proper conditions, the process may be reversed. The same quantity of water may pass through this cycle of changes an indefinite number of times, but it will still be water. If the steam shown in Fig. 28 instead of being cooled in the condenser is passed over certain heated metals, a marked change takes place, as can be easily demonstrated by the following experiment:

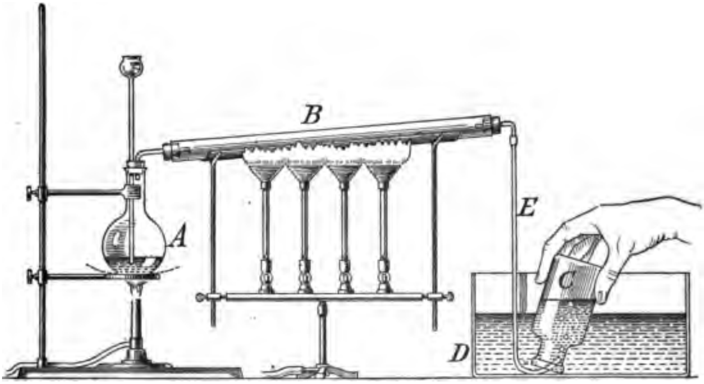


FIG. 44. — Apparatus for the decomposition of water.

A small quantity of powdered zinc is placed in the hard glass tube *B* in Fig. 44. The flask *A* is one third full of water and the bottle *C* is filled with water and inverted in the dish *D*. The water in *A* is boiled until the steam has driven all the air out of the apparatus, and then the burners are lighted under the hard glass tubing B. When the zinc becomes hot the end of the rubber tubing E is slipped under the mouth of the bottle C. Bubbles of gas will soon enter the bottle. When the bottle is filled with the gas, the rubber tubing is withdrawn and the flames are extinguished.

The gas in the bottle is colorless and as far as outward appearance goes might be air. It cannot be steam; for if it were, it would be cooled by the water and condensed. If the bottle is carefully lifted mouth down and a lighted splint or candle is introduced, *the gas will ignite and burn with a pale blue flame.* Clearly here is a gas which has been obtained from the water but which has none of the characteristics of water. The heated zinc has decomposed the water, and this gas is one of the resulting products. This gas was first discovered by the English investigator Cavendish in 1766. It was named *hydrogen* by the French chemist, Lavoisier, because it is one of the constituents of water.

EXERCISES

Ex. **12.** Add one gram of salt to a test tube of water. Do the same with pure sand. What difference do you notice? Are all substances soluble? Try a little gypsum (calcium sulphate). Can you prove that any of it dissolves? Are all soluble substances equal in solubility?

Ex. **13.** Add 1 cc. of alcohol to 10 cc. of water. Do the same with coal oil. What difference do you notice? What is meant by the statement that a liquid is miscible with water?

Ex. **14.** Heat some well water or tap water in a beaker. What are the bubbles that appear? Are gases soluble in water? How does pressure affect the solubility of gases in water? How does rise of temperature affect the solubility of gases? Of solids?

Ex. **15.** Dissolve ten grams of salt in 100 cc. of water in a flask and heat the solution to boiling. Determine the temperature of the solution. How does the dissolved salt affect the boiling point? Wipe off the thermometer and test the temperature of the steam above the water. What is the temperature? To fix the 100° point in a thermometer, would it be best to put the bulb in the boiling water or in the steam? If salt were placed in the water in the outer pan of a double boiler, what effect would it have on the temperature of the material being cooked in the inner pan?

Ex. **16.** Mix one gram of bicarbonate of soda with a like amount of dry powdered tartaric acid. Do you notice any change? Add a few drops of water. What change do you notice now? What effect does water have on chemical action?

Ex. **17.** What is meant by a potable water? Is the mineral matter in well water ordinarily injurious? What is the chief source of danger in drinking water? Why should wells never be near cesspools? Which are safer, deep or shallow wells? How should wells be protected at the surface? How can contaminated water be made safe to drink? Is cistern water ever fit for drinking?

Draw a plan, where well or spring water is used, showing the location of your water supply and its distance from cesspools, manure piles, or other possible sources of sewage contamination. Observe also whether the natural drainage slope from such sources of contamination is towards or away from the water supply. Do you find that a safe or an unsafe condition exists? Can unsatisfactory conditions be improved by change in location, or in a modification of the direction of drainage? Draft a plan which if applied would be an improvement on the present arrangement. What is the only safe course to pursue with reference to drinking water when changes in present conditions cannot be made?

Ex. **18.** Make fine shavings of Castile soap and dissolve .5 gram in 50 cc. of alcohol. Place 25 cc. of rain water in an 8-oz. bottle. By means of a medicine dropper add the soap solution, one drop at a time, to the water and shake vigorously after each addition. Continue until a suds is formed which will stand for at least a minute. Record the number of drops of soap solution used. Repeat the experiment with distilled water, well water, and boiled well water. Which water required the most soap? Which the least? Why is rain water preferred to well water in the laundry? Was there any difference between the boiled and the unboiled well water? What is meant by hardness of water? By temporary hardness? By permanent hardness? Examine the inside of the teakettle at home. Is there any mineral matter on the inside of the kettle? How do you account for it? Would you expect more scale in the kettle in a sandstone country or in a limestone country? What are mineral waters?

Ex. **19.** Heat small quantities of grass, potato, or other vegetables in a test tube (Fig. 43). What do you notice? Repeat the experiment with a bit of lean meat. What can you say about water in organic matter?

Ex. **20.** Select a clear crystal of sodium sulphate and heat in a test tube. Does water come off? What is this water called? What has happened to the crystal? Repeat the experiment with a crystal of copper sulphate. What happens to this crystal? Has the color changed? Add a few drops of water. What happens to the color? Dissolve in a small quantity of hot water and set aside to crystallize. Do the crystals resemble the one you started with? Is water necessary to all crystals? Heat a crystal of potassium bichromate. What do you observe in this case?

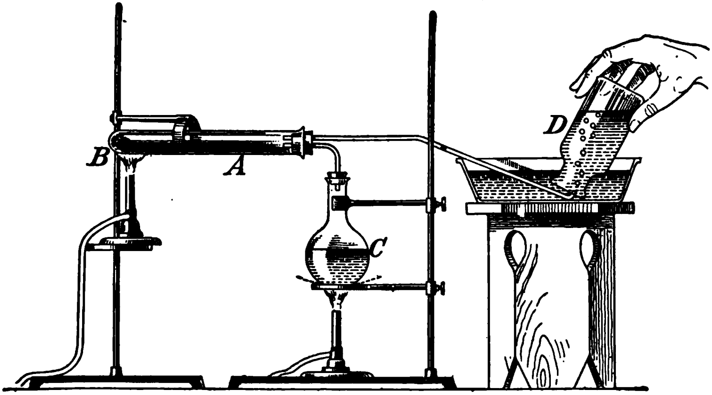


FIG. 45. — **A simple apparatus for decomposing water by means of zinc dust.**

Ex. **21.** Select a bright crystal of washing soda and allow it to stand exposed to the air for several hours. What happens to the crystal? What term is used for such substances? Try the same experiment with calcium chloride. What do you observe in this case? What term is used for such substances?

Ex. **22**. Arrange apparatus as in Fig. 44. The tube *B* may be of hard glass or it may be a piece of one-half-inch iron gas pipe. Place powdered zinc in the tube. Boil the water in the flask *A* so that steam will pass over the zinc, which is heated to low redness by the burner beneath. The air is first driven out, and then pure steam passes over. Finally when the zinc becomes hot gas begins to appear. Collect one or two bottles of gas in the manner described in Section 21. What is the appearance of the gas? Carefully lift the bottle and apply a lighted splint to the mouth. What happens? What is the name of this gas? Who first discovered it?

Note. This experiment may be conducted in the apparatus shown in Fig. 45. *A* is a hard glass test tube. Zinc dust is placed at *B*. The steam from *C* is decomposed in the same way as in Ex. 22. This experiment is made more striking if magnesium tape is used in the place of the zinc, as the magnesium can be seen to burn in the steam. A piece of magnesium tape one foot long will furnish sufficient gas to fill an 8-oz. bottle.

CHAPTER III

HYDROGEN

22. Preparation. Hydrogen may be prepared by passing steam over heated zinc, iron, or other metals in the manner described in the preceding chapter, but these methods are not convenient when the gas is required in large quantities. The preparation of hydrogen in the laboratory is usually accomplished by the action of sulphuric acid on zinc or iron. Zinc is ordinarily employed because most samples of iron contain impurities which contaminate the hydrogen. In the wide-mouth bottle *A* (Fig. 46) there is placed a small quantity of granulated zinc, and water is poured into the thistle tube *B* until the zinc is covered. The end of the thistle tube must be beneath the water. Dilute sulphuric acid is then poured down the thistle tube. An active evolution

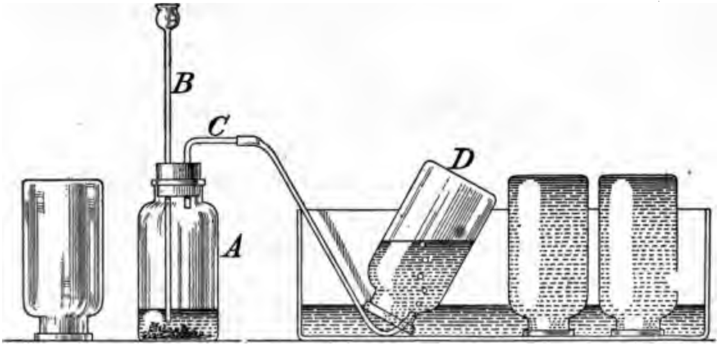


FIG. 46. — Apparatus for the preparation of hydrogen.

of gas takes place, and when sufficient time has elapsed for the hydrogen to drive all the air out of the apparatus, the gas is collected over water in the wide-mouth bottle *D*. The hydrogen in this case is contained in the sulphuric acid and is replaced or driven out by the zinc.

23. Properties. Hydrogen is a gas that has no taste, color, or odor. It will burn readily, and when it is mixed with air the combustion takes place with an explosion. Hydrogen is the lightest known substance and weighs less than one fourteenth as much as air. For this reason it was used formerly to fill balloons; but it is now largely superseded by ordinary illuminating gas, which is less expensive. It is still used for the airship or dirigible balloon. That hydrogen is lighter than air is shown by the fact that it can be poured upwards. A bottle full of hydrogen is gradually turned mouth upward beneath an inverted bottle filled with air (Fig. 47). If after a minute or two a lighted splint is applied to the mouths of the two bottles, it will be found that the hydrogen has passed out of the bottle *A* into *B*. The same thing is shown by blowing soap bubbles with hydrogen; in which case, owing to the lightness of the gas, the bubbles will rise rapidly in the air.

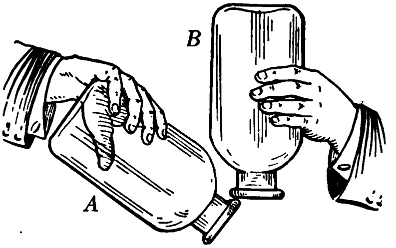


FIG. 47. — Pouring hydrogen upwards.

24. When Hydrogen Burns Water is Formed. The apparatus shown below may be used to study the behavior of hydrogen when burning quietly in the air (Fig. 48).

The bottle *A* is the hydrogen generator already described.

B is a tube containing calcium chloride to dry the hydrogen. After the gas has been carefully tested to make sure that all air has been driven out of the apparatus, the jet is ignited at *C*. The flame will be found to be intensely hot. Pure hydrogen burns with a pale blue flame, but in such an apparatus it is likely to be colored yellow by the sodium in the glass. If a bell jar or a large bottle is held in the position shown by *D*, moisture will collect on the interior until it runs down the sides as indicated. *When hydrogen is burned, water is always produced.*

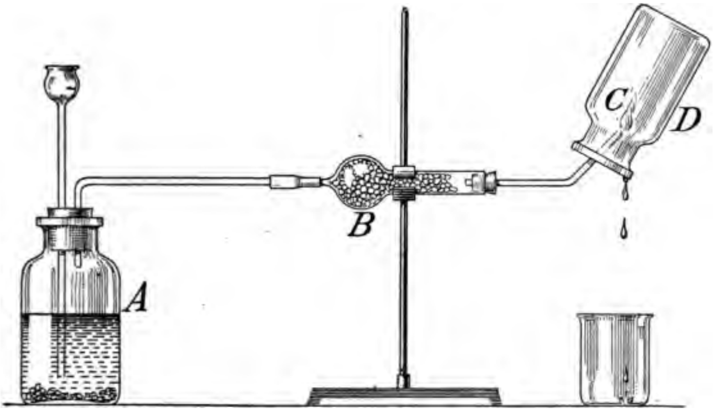


Fig. 48. — **Apparatus for showing behavior of hydrogen when burning in the air.**

25. Occurrence. Hydrogen is not found in nature in the free state except in the merest traces. It is, however, very abundant and very widely distributed in combination with other substances. It is found in water as has been shown, and it is also a part of nearly all animal and vegetable substances. Crude petroleum and all the products made therefrom contain hydrogen. When any substance containing hydrogen is burned in air, *water is formed,* as can be shown by holding a cold bottle or piece of glass over the flame of a candle, kerosene lamp, or gas jet. Candles, kerosene, and gas all contain hydrogen, and a film of moisture will be deposited on any cold surface held above the flame.

26. Water Contains Something Besides Hydrogen. In the experiment described in section 21, it was shown that hydrogen could be prepared from water. This gas is so different from water, however, that water must contain something in addition to hydrogen; or, in other words, hydrogen is combined with something else to form water. An examination of the zinc in the hard glass tube will show that part of it has changed in appearance, an occurrence which indicates that something from the water has combined with the zinc. It will be interesting to try to ascertain what the other substance or substances in water may be.

27. Electrolysis of Water. Water and many other substances can be decomposed by the electric current. The effect of the electric current on water can be readily shown in the apparatus illustrated in Fig. 49. Platinum wires, to the ends of which are attached pieces of platinum foil, are fused into the tubes *B* and *C*. The stopcocks at the top of the tubes are opened, and the apparatus is filled with water containing about one tenth of its volume of sulphuric acid. The water is put into the apparatus at *A*. The sulphuric acid is used because pure water will not conduct electricity. The stopcocks are now closed, and the platinum wires in the tubes *B* and *C* are connected with wires leading to a battery. Three or more dry cells, or two or three dichromate cells, will serve the purpose. As soon as the current begins to pass through the water, bubbles of gas will be noticed passing from the platinum foil to the upper part of the tube. It will be seen that more gas collects in one tube than in the other, and it will soon be apparent that one tube contains exactly twice as much gas as the other. Both of the gases must have come from water, for a careful analysis would show that all the sulphuric acid added still remains in the apparatus.

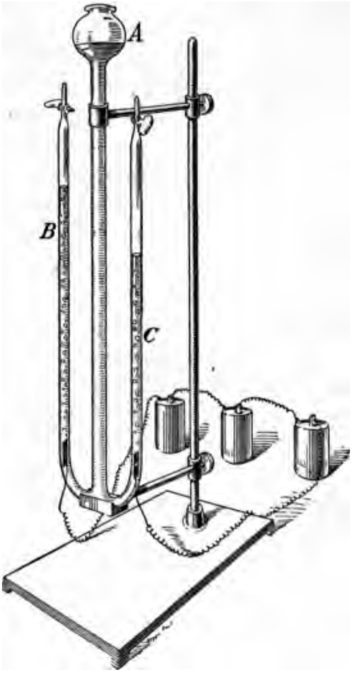


FIG. 49. — **Method of decomposing water by the electric current.**

The larger volume is on the side connected with the negative pole of the battery. When this tube is nearly full of the gas, the battery is disconnected and the gases are examined. By attaching a small rubber tube to the tip of the stopcock the gases may be collected in test tubes over water in the usual way. The gas from the side containing the double volume burns in such a way as to show that it is hydrogen, but the other gas behaves quite differently. It does not burn, but if a splint with a glowing coal on the end is thrust into the test tube containing this gas, *it will burst into a flame.* Here, then, is a new gas which does not burn, but which causes the splint to burn more vigorously than it did in the air.

This gas was discovered at about the same time (1774–75) by the English chemist Priestley and the Swedish chemist Scheele, although each was working independently. Lavoisier, the French chemist, gave it the name *oxygen* (meaning acid-former), because he thought that all acids owed their properties to this substance, a view now known to be incorrect.

EXERCISES

Ex. **23.** Arrange apparatus as in Fig. 46. *A* is an 8-oz. wide-mouth bottle. Place 10 grams of granulated zinc in this bottle and add sufficient water to cover the zinc and the end of the thistle tube. Prepare dilute sulphuric acid by slowly pouring 15 cc. of the strong acid into 50 cc. of water, stirring constantly. Never pour the water into the acid. Now pour a little of the dilute acid down the thistle tube.

(*Note.* Chemically pure zinc and sulphuric acid will not react. It is advisable, therefore, to add a few drops of a solution of copper sulphate to the water in the bottle *A*.)

Acid is added from time to time to keep up a steady flow of gas. After the air has all been driven out of the apparatus collect four bottles of the gas. What was the source of the hydrogen? Describe the appearance of the gas.

Ex. **24.** Hold a bottle of the gas mouth downward and apply a lighted splint. Describe the result. What does this prove about the gas?

Ex. **25.** Fill a 2-oz. wide-mouth bottle half full of water, invert in the pan of water, and fill with hydrogen. What does the bottle contain? Hold it at arm's length, mouth downward, and ignite over a candle or burner. What happens? What does this prove?

Ex. **26.** Gradually turn a bottle of hydrogen mouth upward under a similar bottle filled with air (Fig. 47). After a minute apply a lighted splint to both bottles. Describe the results. Is hydrogen lighter or heavier than air?

Ex. **27.** Arrange apparatus as in Fig. 48. Add the dilute sulphuric acid, and when hydrogen has been evolved for some minutes test the purity of the gas by collecting test tubes of the gas and igniting them, mouth downward. If all air is out of the apparatus, the hydrogen will burn quietly; otherwise it will explode. When you are sure that only pure hydrogen is coming off, wrap a towel around the bottle and ignite at *C*. Test the heat of the flame by holding a piece of iron in it. Hold a large wide-mouth bottle or bell jar over the flame. What is the result? What is formed when hydrogen burns?

Ex. **28**. Hold a wide-mouth bottle or a bell jar over the flame of a candle, an alcohol lamp, or a gas burner, at home. Does moisture collect on the side of the bottle? What is the source of the moisture?

Ex. **29**. Fill the apparatus shown in Fig. 49 with water containing one tenth its volume of sulphuric acid. Connect the platinum wires with a battery (four dry cells will do) and watch the gas collect in the two arms of the apparatus. Which side contains the most gas? With which pole of the battery is it connected? What is the source of these two gases? When the tube containing the larger quantity of gas is nearly full, disconnect the battery. Place a piece of small rubber tubing over the tip of the stopcock on the tube containing the double quantity of gas and collect the gas in a test tube over water. Test it with a flame. What is this gas? Collect the gas from the other side in the same way. Test it by holding the test tube mouth upward and thrusting into it a splint with a glowing coal on the end. What happens in this case? Does this gas behave like hydrogen? What name has been given to this gas? Who discovered it? By whom was it named?

CHAPTER IV

OXYGEN

28. Preparation. Oxygen may be prepared from water by electrolysis, as described in the last chapter, but it is more commonly prepared in the laboratory by heating some substance which contains it. The material generally used is the white salt known as potassium chlorate (often called chlorate of potash). This substance contains thirty-nine per cent of oxygen, which may be driven off at high temperature. For some unknown reason, however, the oxygen is liberated at a much lower temperature if the potassium chlorate is mixed with about one fourth its own weight of the black substance known as manganese dioxide. The manner of preparing oxygen is illustrated in Fig. 50.

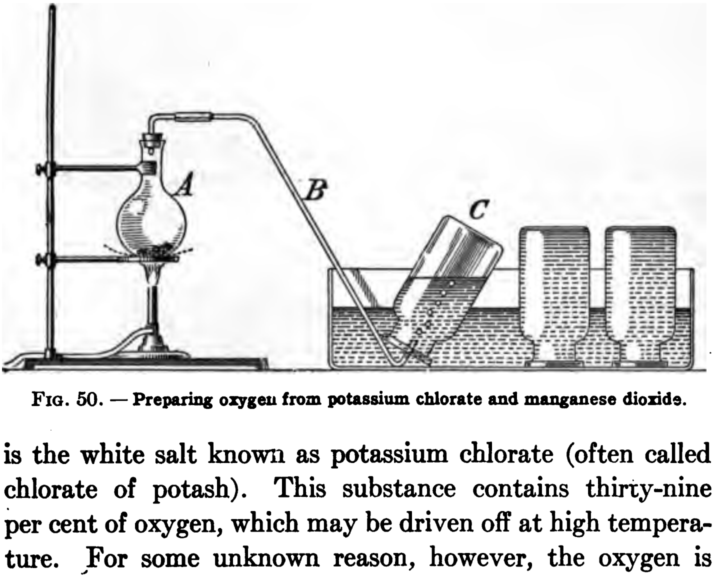


FIG. 50. — **Preparing oxygen from potassium chlorate and manganese dioxide.**

The mixture of potassium chlorate and manganese dioxide is placed in the flask *A*, which stands on a small sandbath. When this is gently heated, the gas soon begins to pass through the tube *B* and is collected over water in the bottles *C*. The gas is not collected until the air has all been driven out of the apparatus. Before the burner is removed (at the end of the experiment) the rubber tube *B* is disconnected to prevent the water from being drawn back into the flask as it cools. When large quantities of oxygen are to be prepared, the copper flask, shown in Fig. 51, is used instead of a glass flask, which is easily broken.

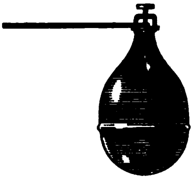


FIG. 51. — **Copper flask for making oxygen.**

29. Properties. Oxygen is a colorless, odorless, and tasteless gas, which is 1.1 times heavier than air. Its most interesting property is the way in which it supports combustion. Materials which burn in the air burn much more rapidly in oxygen. If a splint with a glowing coal on the end is thrust into a bottle of oxygen, it will burst into flame and burn vigorously. Sulphur burns with a feeble flame in the air, but in oxygen the flame is increased in size and brightness. Phosphorus burns readily in the air, but in oxygen it burns with a dazzling brilliance. Some substances that will not burn in the air will do so in pure oxygen. Take a piece of picture frame wire, heat the end in a Bunsen burner, and dip into powdered sulphur. The sulphur will adhere to the end of the wire, and if it is ignited in the burner and then thrust into a jar of oxygen, the wire will

FIG. 52. — **Burning iron wire in oxygen.**

burn with brilliant scintillation (Fig. 52). Glowing balls of molten matter drop from the wire to the bottom of the jar.

Hydrogen mixed with half its volume of oxygen explodes when ignited much more violently than it does when mixed with air. Experiments with mixtures of these two gases should be conducted with great caution. The flame made by burning hydrogen in oxygen is, with one exception, the hottest known. To make use of this flame a special oxy-hydrogen burner is used which permits the gases to mix only at the nozzle (Fig. 53). When this flame plays against a piece of lime, it heats it to a white heat and forms the so-called " limelight," sometimes used for stereopticons.

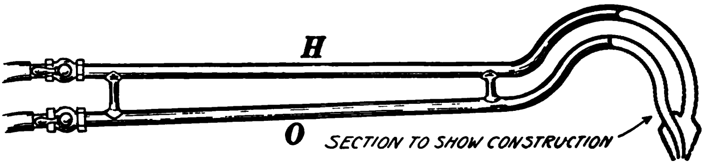


FIG. 53. — Oxy-hydrogen blowpipe.

30. Composition of Water. The decomposition of water by electrolysis shows that both hydrogen and oxygen are contained in water, but this experiment alone does not prove that water is composed *solely* of hydrogen and oxygen. If these two substances can be combined to form water, the proof that water is composed only of these two substances will be conclusive. This has frequently been done by means of the apparatus shown in Fig. 54.

A is a graduated tube, known as a *eudiometer*, which has two platinum wires fused into the closed end, the ends of the wires being about 2 mm. apart. This tube is filled with mercury and inverted in a cylinder *B* full of mercury. Hydrogen gas is then introduced into the tube *A*, and the volume is carefully determined by raising or lowering the eudiometer until the mercury is at the same level inside and outside of the tube. It is an easy matter to determine the volume since it may be read on the scale. Oxygen gas is then introduced and measured in the same way. Nothing occurs to the mixture of gases if allowed to stand at ordinary temperatures; but if an electric spark is caused to pass between the two platinum wires, the gases will unite with a slight explosion. If the gases are mixed in the proportion of *exactly two volumes of hydrogen to one of oxygen,* they will disappear entirely, and nothing will be left but a minute quantity of water. That the quantity of water is so small is due to the fact that the combined volumes of the hydrogen and oxygen required to produce water is over 2500 times the volume of the resulting water. For example, if one cubic centimeter of water were decomposed into hydrogen and oxygen, the combined gases would have a volume of over 2500 cubic centimeters.

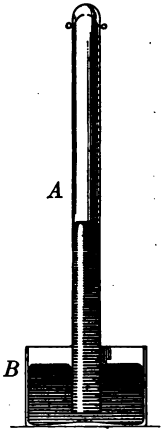


FIG. 54. — Eudiometer.

31. Proportion of Hydrogen and Oxygen in Water. If the hydrogen in the eudiometer is more than twice the volume of the oxygen, an excess of hydrogen will remain after the explosion. On the other hand if the oxygen is more than half the volume of the hydrogen, an excess of oxygen will remain uncombined. Repeated experiments of this kind have demonstrated that water is composed of hydrogen and oxygen, and that these gases always combine in the ratio of one volume of oxygen to two of hydrogen. As the weight of a given volume of oxygen is sixteen times that of the same volume of hydrogen, it follows that water is composed of one part by weight of hydrogen to eight parts by weight of oxygen.

32. Analysis and Synthesis. Two different methods have been suggested for determining the composition of water. In the last chapter the method described was the decomposition of water by means of the electric current. A process in which a substance is separated into its component parts is termed *analysis*. The decomposition of water by means of heated zinc resulting in the formation of hydrogen is also a method of analysis. Such a process as the formation of water in the eudiometer, in which a substance is formed by bringing about the combination of its component parts, is termed *synthesis*.

33. Elements and Compounds. It has been shown that water, which appears to be a simple substance, is in reality composed of two substances — hydrogen and oxygen. It is natural to think that if water can be decomposed into two substances it might be possible to decompose hydrogen and oxygen. All attempts to get anything else out of these substances, however, have failed. In other words, hydrogen and oxygen are such simple substances that so far as is now known they cannot possibly be divided any further. Such substances which cannot be split up into anything simpler are called *elements*.

Water, which is a union of hydrogen and oxygen, is a representative of that class of substances which are composed of two or more elements united. Substances that are composed of two or more elements are called *compounds*. Water, therefore, is a compound.

The study of chemistry has shown that the number of elements is quite small, only about eighty in fact. Of the known elements, not more than twenty are of everyday importance. These, only, will be studied in this text. A complete list of the elements is given on the inside of the back cover.

Although the number of elements is small, many thousands of compounds are known, and the list is being continually increased. Water, sugar, salt, alcohol, ammonia, and starch are familiar compounds. Oxygen, iron, lead, sulphur, and carbon are well-known elements.

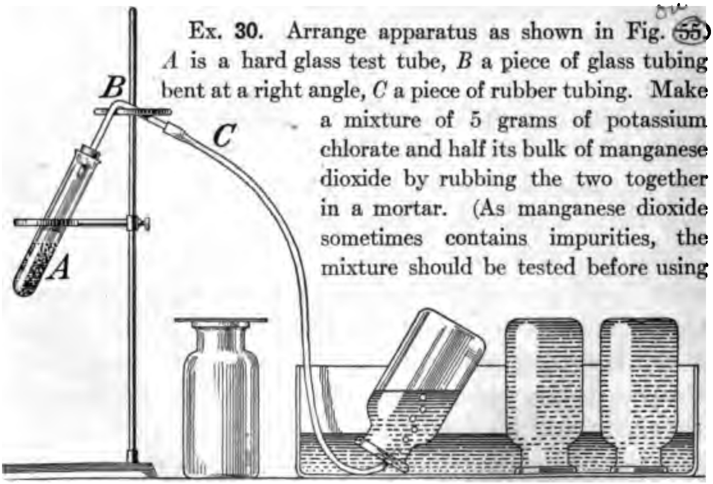
34. Definite Proportions. It has been shown (31) that water is composed of eight parts by weight of oxygen to one part by weight of hydrogen. This proportion never varies no matter whether the water be obtained by melting pure ice, by condensing steam, or by burning hydrogen. Careful analysis of many chemical compounds has demonstrated that this is a general rule, and that *a chemical compound always contains the same elements,* and that these elements are *always present in exactly the same proportions by weight.* This fact is known as the *law of definite proportions.*

35. Physical and Chemical Changes. Two different kinds of changes in water have also been observed. Ice can be changed to water and water to steam; and the process may be reversed. But whether it exists as a solid, liquid, or vapor, it is still water. Such a change, which merely affects the form and does not affect the composition of the substance, is known as a *physical change.* When water is decomposed into hydrogen and oxygen, or when these two gases are united by the heat of the electric spark, the change which takes place is quite different; for in either case the products of the change are entirely different from the substance or substances undergoing the change. Such a change as this, which affects the composition of the substance, is called a *chemical change.*

Two other terms may be explained here; namely, *mechanical mixture* and *chemical compound.* In a mechanical mixture the substances do not undergo a chemical change, but each retains all of its individual characteristics. The mere mixing of the gases hydrogen and oxygen in the eudiometer is a good example. In the case of a *chemical compound* two or more elements have undergone a chemical change, and have united and thus lost their original characteristics. Water, which is formed by the chemical union of hydrogen and oxygen, is a good example of a chemical compound.

Physical changes and mechanical mixtures may be said to be in the realm of *physics,* while the chemical changes are in the field of *chemistry.*

EXERCISES



Ex. 30. Arrange apparatus as shown in Fig. 55. *A* is a hard glass test tube, *B* a piece of glass tubing bent at a right angle, *C* a piece of rubber tubing. Make a mixture of 5 grams of potassium chlorate and half its bulk of manganese dioxide by rubbing the two together in a mortar. (As manganese dioxide sometimes contains impurities, the mixture should be tested before using

FIG. 55. — A simple apparatus for generating oxygen.

by heating about one gram in a test tube. If the oxygen comes off quietly, proceed with the experiment.) Place the mixture in the test tube and heat gradually, beginning at the top. A steady flow of gas may be obtained by regulating the amount of heat applied. Collect four 8-ounce bottles of the gas in the usual way. At the end of the experiment withdraw the tubing from the water, or remove the cork from the test tube. What is the appearance of the gas? (The cloudiness which sometimes shows at the beginning of the experiment will soon disappear.)

Ex. **31.** (*a*) Slip a glass plate over the mouth of one of the bottles and turn it mouth upward. Thrust a splint with a glowing coal on the end into the bottle repeatedly.

(*b*) Warm a glass rod so that a piece of sulphur (the size of a grain of wheat) will adhere to it. Ignite the sulphur and note how it burns. Thrust it into a bottle of oxygen and note whether it burns differently.

(*c*) Take a piece of picture frame wire 8 inches long and spread one end slightly. Heat this end of the wire and dip it into powdered sulphur. Ignite the sulphur which adheres to the wire and thrust it into a bottle of oxygen. What happens to the glowing coal? Does the oxygen ignite as hydrogen did? What difference do you note in the sulphur burning in air and oxygen? Is the odor the same in each case? Do substances that burn in air burn more or less vigorously in oxygen? What happens in the case of the iron wire? Will the wire burn in the air? What can you say in general about burning in oxygen as compared with burning in air?

Ex. **32.** (By the Teacher.) Hollow the end of a piece of crayon into a small cup and attach a piece of wire as shown in Fig. 56. Place a piece of phosphorus the size of a grain of wheat in the cup and ignite by touching it with a hot wire. Lower it into a bottle of oxygen and note results. (Phosphorus must be handled with extreme care. It is always kept covered with water and should always be cut under water. Never touch phosphorus with the hands, but handle with forceps.) How does the burning in air compare with burning in oxygen? What is the source of the white fumes? Do the white fumes dissolve in water?

FIG. 56. — A crayon spoon for burning phosphorus.

Ex. **33.** (By the Teacher.) Fill a half pint cream bottle one third full of oxygen and two thirds full of hydrogen. Wrap the bottle in a thick towel and holding at arm's length quickly bring it mouth downward over a flame. How does the explosion compare with that of hydrogen and air? What is meant by the oxy-hydrogen flame? How does it compare with other flames for heat? Draw a diagram of an oxy-hydrogen burner. How is the limelight obtained?

Ex. **34.** What is analysis? Synthesis? What is an element? A compound? Give examples of each. How does the number of elements compare with the number of compounds? What is a physical change? A chemical change? A mechanical mixture? A chemical compound? With what kind of changes is chemistry concerned?

CHAPTER V

OXYGEN (*Continued*)

36. Oxides and Oxidation. The experiments with oxygen described in the last chapter show that at ordinary temperatures oxygen has little effect upon substances placed in it, but that at higher temperatures it makes them burn more actively than they will in the air. Oxygen is said, therefore, to be rather *inactive* at low temperatures, but very *active* at high temperatures. What really happens, when the substances burn in oxygen, is that the oxygen combines with the burning materials. Oxygen unites with sulphur and forms a suffocating gas; with phosphorus it forms white fumes; and with iron it forms the black molten material which drops to the bottom of the bottle.

When oxygen unites with another element, the result is a compound called an *oxide*. Thus iron and oxygen form iron oxide; phosphorus and oxygen form phosphorus oxide. Some of the oxides are colorless gases like the oxides of sulphur and carbon, but the majority of them are solids, such as the oxides of iron, phosphorus, and lead. When oxygen unites with another element, the process is called *oxidation*, and the resulting compounds formed are known as *products of oxidation*. Water is the product of the oxidation of hydrogen, and is in reality *hydrogen oxide*.

37. Burning in Air due to Oxygen. The strong resemblance between the burning of substances in oxygen and in the air suggests that these two processes are the same. Moreover, carefully conducted experiments show that whether the substance is burned in oxygen or in the air the products formed are exactly the same. It is certain, therefore, that the process of burning in the air is due to the presence of oxygen. The fact that substances burn less readily in air than in oxygen further suggests that the air is not pure oxygen. The proportion of oxygen in the air can be shown by a simple experiment (Fig. 57).

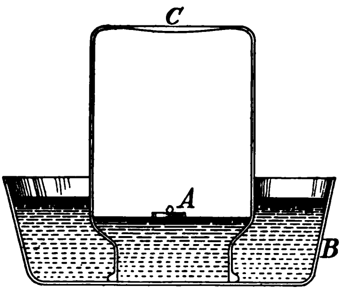


FIG. 57. — **Experiment to show proportion of oxygen in the air.**

A small piece of cork or of wood *A* is floated in the water in the pan *B*. A piece of phosphorus the size of a pea is placed on the cork. The phosphorus is ignited by a touch from a piece of hot iron wire, and a wide-mouth bottle *C* is carefully placed over it. The phosphorus will burn until the oxygen in the air within the jar is exhausted, and then it will be extinguished. The water rises in the bottle to take the place of the oxygen consumed. The white cloud of oxide of phosphorus will soon be dissolved in the water, and if the bottle is so adjusted that the level of the water inside and outside is the same, it will be found that the jar is about one fifth full of water. The air, therefore, is about one fifth oxygen and four fifths of some other gas which does not unite with the phosphorus.

38. Combustion Defined. Combustion, in its ordinary sense, whether in oxygen or in the air, consists in the union of substances with oxygen with the evolution of light and heat.[1] Substances which will unite with oxygen are said to be combustible and those which will not are incombustible. All the important elements will unite with oxygen.

39. Kindling Temperature. It has been shown that substances do not usually combine with oxygen at ordinary temperatures. This is a fortunate thing; for if it were not so, all combustible substances in nature would burn up, since there is sufficient oxygen in the air for that purpose. Some substances need to be heated very little before they will burn, while others ignite only when raised to very high temperatures. If small pieces of phosphorus, sulphur, and charcoal are placed on an iron plate with a lighted burner beneath, the phosphorus soon bursts into a flame. The sulphur, however, requires considerable heat to ignite it, and the charcoal does not burn until it reaches red heat. Similar experiments with a variety of substances demonstrate that each has a certain temperature to which it must be heated before it will ignite. This is known as the *kindling temperature* of the substance. When the material begins to burn, the heat of the burning parts raises the temperature of the adjacent parts to the kindling temperature and the burning spreads. The way the flame creeps

FIG. 58. — The difference in the kindling temperature of paper, wood, and coal is utilized in building a coal fire.

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[1] Chemists sometimes define combustion as "any rapid chemical action accompanied by light and heat." There is a limited number of such chemical actions in which oxygen takes no part.

along a stick of wood is an example in point. Practical advantage is taken of the difference in kindling temperatures in lighting a coal fire, by the fact that paper is first placed on the grate bars, then pine kindling is placed on the paper, and then the coal above (Fig. 58). The paper is easily ignited, but since it alone cannot raise the coal to its kindling temperature, the pine is placed between.

40. Slow Oxidation. Some substances unite slowly with oxygen at ordinary temperatures without the evolution of light. Iron is such a substance. When iron rusts, as it does when exposed to moisture, the change consists in the union of the iron with oxygen, and the rust is similar to the substance formed when iron is burned in oxygen. If some moist iron dust or filings are placed on the floating cork shown in Fig. 57, and the apparatus allowed to stand for one or two weeks, the water will be seen to rise slowly in the bell jar. If allowed to stand long enough, it will be found that, as in the case of the burning phosphorus, one fifth of the volume of the bottle is filled with water. A change of this kind is known as *slow oxidation*. Slow oxidation is of common occurrence. It is always taking place in the animal body. The oxygen taken into the lungs acts upon various substances in the body, oxidizing them into other forms which can be more readily eliminated, as, for example, water and an oxide of carbon. The decaying of wood and other vegetable and animal matter is a slow oxidation process brought about by the action of bacteria.

What difference is there between combustion and slow oxidation? Apparently in the case of the latter there is no light or heat produced. A careful study of the subject has shown, however, that a given amount of iron produces *exactly the same amount of heat* whether burned in oxygen or allowed to rust in the air. In the one case the heat is all given off in a short time and the temperature becomes so high that light is emitted. In the other case the heat is evolved slowly, and the surrounding air conducts it away as rapidly as it is produced. *The quantity of heat is the same in both cases.*

41. Weight Relations of Combustion. If the burning of a substance consists in its union with oxygen, it follows that the products of combustion must weigh more than the substance burned. The oxide resulting from the burning of the iron wire weighs more than the wire itself. The oxide produced when copper was heated in oxygen is heavier than the copper burned. In the case of burning wood, or a lighted candle, there is apparently a loss of matter, for the wood and candle almost completely disappear. The wood and the candle are composed largely of carbon and hydrogen. When the latter burns, it forms the oxide of hydrogen called water and disappears into the air as vapor; but it has been shown that the water produced weighs nine times as much as the hydrogen burned (31). The carbon burns to an oxide which is a colorless gas, and it also disappears into the atmosphere. This gas weighs nearly four times as much as the carbon burned. Figure 59 shows a method of demonstrating that the products of combus-

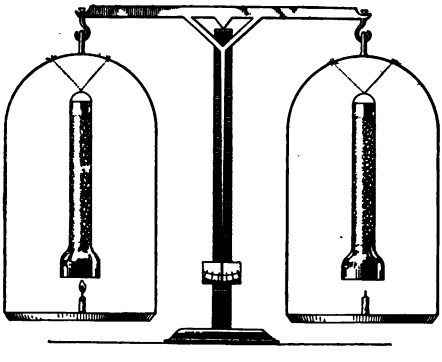

FIG. 59. — Experiment to show the weight relations of combustion.

tion of a candle weigh more than the candle itself. Two lamp chimneys are suspended on each side of a balance over unlighted candles. The upper parts of the chimneys are filled with caustic soda (sodium hydroxide), which has the power of absorbing water and the oxide of carbon. The pans are exactly balanced by placing sand or small weights on the lighter side. One of the candles is then lighted, and as it burns away and the products of combustion are absorbed by the caustic soda in the chimney above the burnt candle, that side of the balance slowly sinks, showing that it has increased in weight.

This experiment shows that the elements in the candle were not actually destroyed, but merely changed into other compounds by uniting with oxygen. In fact, scientists believe that matter cannot be destroyed, and that after every physical or chemical change the amount of matter is the same as before. This is merely the equivalent of saying that it is not possible either to create or destroy matter. This is known as *the law of the indestructibility of matter,* sometimes known as the law of the *conservation of matter.*

42. Spontaneous Combustion. While slow oxidation usually takes place in such a way that there is no perceptible rise of temperature, it is possible for the oxidizing substance to be so placed that the heat cannot radiate as fast as formed. In that case the heat may accumulate until it has warmed the material to its kindling temperature, upon reaching which it will burn. This is probably the cause of the so-called *spontaneous combustion.* Some oils, notably linseed oil, oxidize readily, and oily rags have often been discovered to be on fire. Barns have been burned by the spontaneous combustion of hay, and fires that are to be accounted for only by spontaneous combustion have been found in the center of large heaps of coal.

43. Reduction the Opposite of Oxidation. It has been said that the union of oxygen with other elements is called oxidation; but when hydrogen is passed over copper oxide the opposite effect is produced, namely, the oxygen is taken away from the copper. The process, which is the reverse of oxidation, is known as *reduction.* Anything which, like hydrogen, causes another substance to lose oxygen is called a *reducing agent.* In this experiment it will be noticed that the hydrogen is *oxidized* but the copper oxide is *reduced.*

44. Occurrence of Oxygen. Oxygen is the most abundant and most widely distributed of all the elements. It comprises four fifths of the water, one fifth of the air, and one half of the rocks of the earth. It is found in all plant and animal bodies, and it is absolutely essential to all life of both the animal and vegetable kingdom. It dissolves to the extent of three per cent in water, making possible the life of fishes. It is used commercially in a number of ways, and it is useful in the treatment of certain diseases. For commercial purposes it is sold in a compressed form in strong steel cylinders (Fig. 60). It is prepared from potassium chlorate, or from liquid air in the manner to be described in the next chapter.

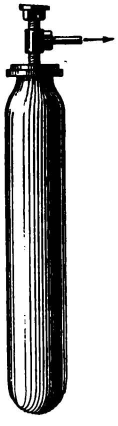

FIG. 60. — A cylinder of commercial oxygen.

45. Hydrogen Peroxide. Water, as has been shown, consists of one part by weight of hydrogen to eight of oxygen. Another compound of hydrogen and oxygen is known which contains twice as much oxygen as water, — that is, 16 parts by weight of oxygen to one of hydrogen. This substance is known as *hydrogen peroxide*. It is an unstable compound and readily gives up half of its oxygen and changes to water. Thus it is a good oxidizing agent. Its use in medicine and in bleaching depends upon this property. It is never used in the pure state. The solution found on the market contains about three per cent of actual hydrogen peroxide dissolved in water.

EXERCISES

Ex. **35**. Is oxygen very active at low temperatures? At high temperatures? What takes place when substances burn in oxygen? When oxygen unites with another element, what is the product called? Name the product of the combustion of oxygen with iron; with phosphorus; with carbon. What is meant by oxidation? What is meant by "products of oxidation"? What is formed when hydrogen is oxidized?

Ex. **36**. (By the Teacher.) Perform the experiment described in paragraph 37.[1] What causes the water to rise in the bottle? What proportion of the air is oxygen? Compare the burning in air with burning in oxygen. Are the products of combustion the same in each case? In which does the burning take place more vigorously? What is meant by combustion as the term is ordinarily used? What is meant by combustible and incombustible substances?

Ex. **37**. Perform the experiment described in paragraph 39. What happened to the phosphorus, sulphur, and charcoal? Which ignited at the lowest temperature? What is meant by kindling temperature? Do substances usually burn at ordinary temperatures? Why is this a fortunate thing? Is there much difference in the kindling temperatures of substances? Is the kindling temperature constant for each substance? What practical use of the difference in kindling temperatures of substances is made in the home? Light the end of a long splinter of wood. What makes the flame spread along the

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[1] Allow the bottle to remain standing mouth downward in the water for use in a later experiment.

wood? Why does the wood burn more rapidly if held with the burning end downward than if the burning end is at the top?

Ex. **38.** Examine a piece of rusty iron from your home. Of what is the rust composed? How does it compare with the product formed when iron burns in oxygen? What is meant by slow oxidation? Is it of common occurrence? Mention a few instances of slow oxidation. What is the source of the heat in animal bodies? Is heat produced during slow oxidation? How does the quantity of heat compare with that produced by rapid combustion of the same substance?

Ex. **39.** Perform the experiment described in paragraph 41. What becomes of the candle when it burns? Do the products of combustion weigh more or less than the substance burned? When wood or coal is burned in the stove, only a small amount of ash is left; what becomes of the remainder of the fuel? Is any of the matter destroyed? What is the law of the indestructibility of matter? Iron rust weighs more than the iron from which it was made; has matter been created? How is spontaneous combustion explained? Why is it dangerous to allow oily rags to lie around? Do you know of any cases of spontaneous combustion in your vicinity?

Ex. **40.** Arrange apparatus as shown in Fig. 61. *A* is an ordinary hydrogen generator. *B* is a piece of glass tubing with the horizontal arm about 8 inches long. *C* is a hard glass test tube with some fine copper oxide at the closed end. Allow hydrogen to pass through the tube for a few minutes and then continuously heat the copper oxide and note the result. Does the copper oxide change color? Why? What becomes of the oxygen which is removed from the copper oxide? What is meant by reduction? Is hydrogen a reducing agent? Is the copper oxide a reducing agent or an oxidizing agent?

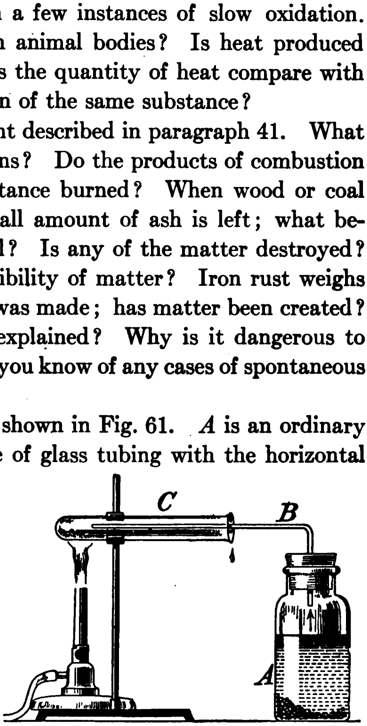

Fig. 61. — **Reducing copper oxide with hydrogen.**

Ex. **41.** Tell what you can about the occurrence of oxygen, (*a*) in the elemental form, (*b*) in combination with other substances. How is it prepared commercially? What is hydrogen peroxide? For what is it used?

CHAPTER VI

AIR — NITROGEN

46. It will be well to return to the experiment in which phosphorus was burned in the wide-mouth bottle (Fig. 57). The bottle is now so adjusted that the level of the water inside and outside is the same, a glass plate is slipped under the mouth, and the bottle turned mouth upward. If, now, a lighted candle is placed in the bottle (Fig. 62), it will not burn as it does in air or oxygen, but will be immediately extinguished. Nor will the gas itself burn as does hydrogen. Here is a new gas, then, which neither burns nor supports combustion. This gas was first studied by the English chemist Rutherford in 1772. It was later named *nitrogen* because of its presence in niter or saltpeter.

Fig. 62. — A candle is extinguished when placed in nitrogen.

47. Properties. Nitrogen is a colorless, odorless, and tasteless gas. It is lighter than oxygen or air and constitutes nearly four fifths of the atmosphere. It is less soluble in water than is oxygen. Chemically it is said to be very inactive, for it will not readily unite with any other element, even at rather high temperatures. It is evident that it does not unite with oxygen at ordinary temperatures. The electric spark will cause a limited union of nitrogen and oxygen, and this combination is supposed to take place to a slight extent during electric storms. It may be said, therefore, that most of its properties are of a negative character, and that nitrogen is most noted in the laboratory for what it will not do. Animals placed in nitrogen die immediately of suffocation, not because the gas is poisonous, but because they cannot live without oxygen.

48. Occurrence. Nitrogen exists as an element in the atmosphere. United with other elements into very complex compounds, it is found in plant and animal tissues. It is sometimes said to be the most important element to life, as it is necessary to the formation of the protoplasm of the cell upon which life depends. While the worker in the laboratory finds difficulty in making nitrogen combine with other elements, nature has a way of bringing about this union. Certain bacteria found in the soil have the power of combining nitrogen with hydrogen and oxygen in such a way that crops can use it. Some of these bacteria grow in little nodules, or tubercles, on the roots of plants like the clovers, peas, beans, and other leguminous plants. Others of these bacteria do not grow on the roots of plants but live free in the soil. Animals eat the crops and get their necessary nitrogen compounds in that way, so that these tiny bacteria are responsible for most of the combined nitrogen found in nature. More will be said about these bacteria later (168).

49. Air a Mechanical Mixture. The ratio between nitrogen and oxygen in the air is so nearly constant that the question might arise whether it is a compound or a mechanical mixture of these two elements. All experiments indicate that the air is merely a mechanical mixture consisting of practically 21 per cent oxygen, 78 per cent nitrogen, and 1 per cent of small quantities of a number of other gases. Two reasons for this belief may be mentioned. First, true chemical compounds do not vary in the least in composition. The variation in the composition of air, though slight, is sufficient to show that it is not a compound. Second, if air is dissolved in water and the air taken out of the water by means of an air pump, it will be found that the ratio of oxygen to nitrogen in this air is as 1 to 2, while in the atmosphere it is as 1 to 4. The ratio of the elements in a true chemical compound does not change when the substance is dissolved in water.

50. Air Contains Water Vapor. The experiment wherein calcium chloride absorbed moisture from the air shows that the atmosphere contains water vapor. This is also demonstrated when moisture collects on a vessel containing cold water. The surrounding air is cooled to such an extent that some of its moisture is deposited on the cold vessel in the form of liquid water (Fig. 63).


FIG. 63. — Moisture collects on the outside of a vessel of cold water.

The amount of moisture in the atmosphere is the most variable of all its constituents and is known as *humidity.* The power of the air to hold water vapor increases with rise of temperature. When air contains all the moisture it will hold, it is said to be *saturated.* At 10° C. one cubic meter of air will hold 9.7 grams of water vapor, while at 20° C. it takes 17.1 grams to saturate it. *Relative humidity,* a term quite commonly used, is the ratio between the amount of water vapor in the air at a given temperature and the amount it could hold when saturated at that temperature. A relative humidity of 80, for instance, means that the air contains 80 per cent of the moisture it is possible for it to hold at that temperature. When the air is saturated, the relative humidity is said to be 100. As air cools, its power to hold moisture decreases, and finally the water vapor condenses as dew. The temperature at which this takes place is known as the *dew point*.

51. Air Contains Carbon Dioxide. If a dish containing limewater is allowed to stand exposed to the air for a short time, a crust of white material forms on the surface. This effect is not produced by pure oxygen, by nitrogen, or by water. If the mouth is placed near a dish of limewater and the breath blown on it, or if a lighted candle is lowered into a wide-mouth bottle containing a little limewater (Fig. 64), the white crust forms very rapidly. Evidently, then, the air contains another gas which is probably identical with some substance found in the breath, and which is also formed when a candle burns. This is a compound of oxygen and carbon known as *carbon dioxide*, commonly called *carbonic acid gas*. This gas will be more fully discussed in the chapter on carbon compounds.

FIG. 64. — A burning candle lowered into a bottle containing limewater.

The amount of carbon dioxide in the air is very small, amounting to only 0.03 per cent or 3 parts in 10,000 in the country, and seldom exceeding 0.06 per cent in the cities where much coal is burned. Small as the quantity is, it is very important; for without it green plants could make no growth. The amount in the air is practically constant; for plants use it at about the same rate at which it is produced by the breathing of animals, by the burning of wood and coal, and by other oxidation processes (107).

52. Traces of Other Substances in Air. In addition to the four substances mentioned, the atmosphere contains small quantities of certain rare gases and traces of sulphur compounds, as well as dust, bacteria, and various substances given off from the lungs.

53. Diffusion of Gases. Before leaving the subject of air it will be well to learn something of those physical properties of gases to which occasional reference must be made. Gases diffuse in all directions regardless of their density. If two bottles (Fig. 65) are placed mouth to mouth, the upper one containing the light gas, hydrogen, and the lower containing air, which is fourteen times as heavy, diffusion begins immediately, and after a few minutes there is hydrogen in the lower bottle and air in the upper. If it were not for this property of gases, the atmosphere would consist of a lower layer of the heavy carbon dioxide, a middle layer of oxygen, and an upper one of nitrogen, which is the lightest of the three gases. Diffusion keeps the gases uniformly mixed, but in the atmosphere this process is accelerated by the mixing action of the winds and other air currents.

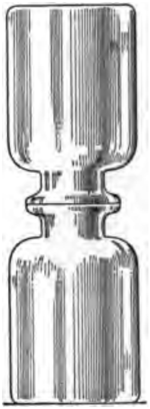
FIG. 65. — To show the diffusion of gases.

54. Effect of Temperature and Pressure. Substances[1] expand when heated, but gases are affected to a much greater extent than liquids or solids. That heated air expands and is, therefore, lighter than cold air is shown by the hot-air balloon so commonly seen at the country fairs. Gases respond readily to change of pressure. The volume of a gas is inversely proportional to the pressure to

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[1] Water below 4° C. is an exception.

which it is subjected. If the pressure is doubled, the volume of the gas becomes one half, provided the temperature remains unchanged. While different liquids and solids expand or contract at varying rates upon change of temperature or pressure, it is interesting to note that *all gases behave the same regardless of their composition.*

55. Liquefaction. All gases can be condensed to liquids and even to solids, provided the proper combination of very low temperature and high pressure can be secured. Oxygen cannot be condensed to a liquid unless its temperature is lowered to − 119° C., while nitrogen must be cooled to − 146° C. and hydrogen to − 241° C. Air can be liquefied in the same way, and when it changes back to the gaseous form the nitrogen boils off first, leaving the oxygen behind. Oxygen of 95 per cent purity is prepared commercially in this way. This behavior of liquid air is another proof that air is a mechanical mixture.

56. Gases Are Substances. That gases are matter and actually occupy space can be shown by a simple experiment. A bottle of air or some other gas is fitted with a tight cork carrying a funnel with a small opening, as in Fig. 66. If the apparatus is air-tight, water poured in the funnel will not run into the bottle, a circumstance which shows that the bottle is already full. If the cork is loosened to allow the gas to escape, the water will enter the bottle.

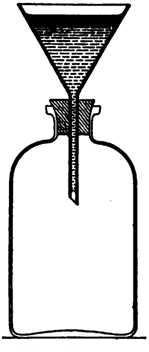


FIG. 66. — The air in the bottle prevents the water in the funnel from entering.

That gases have weight also may be shown by weighing a flask from which the air has been exhausted by an air pump, and then weighing the flask when filled with different gases. A liter (about one quart) of hydrogen weighs 0.09 gram, while a liter of oxygen weighs 1.43 grams.

EXERCISES

Ex. 42. Adjust the bottle left from Ex. 37 so that the level of the water inside and outside of the bottle is the same. Slip a glass plate over the mouth of the bottle and turn it mouth upward. Lower a lighted candle into the bottle. Does the candle continue to burn? Does the gas in the bottle ignite? Is the gas in the bottle different from the two previously studied? What is the name of this gas? Give the properties of nitrogen. Is nitrogen an element or a compound? What can you tell about the occurrence of nitrogen? How does nature cause nitrogen to combine with other elements? How do animals get their nitrogen compounds? Is air a chemical compound or a mechanical mixture? Give reasons for your answer.

Ex. 43. Have you ever noticed moisture gather on the outside of a vessel containing cold water? Where does this moisture come from? In what other way can you demonstrate the presence of water vapor in the air? What is meant by the humidity of air? What is the meaning of relative humidity? When is the relative humidity said to be 100? If air was suddenly heated or cooled, would its relative humidity be changed? Explain. What is meant by the dew point?

Ex. 44. Place limewater in a shallow dish and allow it to stand exposed to the air. What happens to it? Blow the breath into a bottle containing limewater. Lower a candle into a similar bottle. What change takes place in the limewater? Does the air contain something that is also found in the breath and in the products of combustion of the candle? What is this substance? Is much of it present in the air? At home try placing a dish of limewater outdoors and another in the house (in cool weather when the windows are closed) and note if there is any difference in the rate at which the white crust appears in the two dishes. What other substances are found in the air besides nitrogen, oxygen, water vapor, and carbon dioxide?

Ex. 45. Fill an eight-ounce bottle with hydrogen and carefully place it mouth downward over another bottle filled with air as shown in Fig. 65. After letting it stand some time slip a glass plate between the bottles and remove the upper one. Invert the lower bottle and quickly apply a burning splint to the mouth. Have you any evidence that some of the lighter hydrogen has passed downward into the lower bottle? What is meant by diffusion of gases? What effect does this property of gases have on the composition of the atmosphere? Do air currents assist? What effect does pressure have on gases? How are gases affected by changes in temperature? Would a hot-air furnace heat the house as well if placed in the garret instead of in the basement? Why? Should ice be placed in the bottom or in the top of a refrigerator? Why? What makes the hot-air balloon rise? What conditions are necessary to liquefy a gas? When liquid air is boiled, which gas boils off first? Is any practical use made of this fact?

Ex. 46. Fit a funnel with a small opening into a cork and place it in a bottle as in Fig. 66. Pour water into the funnel. Why does the water not run into the bottle? What happens if you loosen the cork?

CHAPTER VII

SULPHUR

57. SULPHUR is a well-known article of commerce in the form of *roll sulphur*, or brimstone, and as the yellow powder called *flowers of sulphur*. It is an element, but unlike those studied so far, is a solid and not a gas. It has been known from the earliest times because it occurs abundantly in nature in the elementary form. It is found in the neighborhood of volcanoes, especially those of Sicily, which

FIG. 67. — **A solid block of Louisiana sulphur.** Molten sulphur is pumped from wells into immense wooden bins, where it solidifies. It is then broken by blasting and loaded on freight cars for shipment.

country was formerly the chief source of the sulphur of commerce. Recently large deposits of sulphur have been discovered in a number of places in the United States. The state of Louisiana (Fig. 67) is the chief producer at the present time.

58. Preparation. Sulphur as found in nature is mixed with earthy matter. If the ore is heated until the sulphur melts, the sulphur may be drawn off in a liquid form, leaving the stones and earth behind. The crude sulphur thus obtained is purified by distillation. The sulphur is distilled into large cooling chambers of brick. When the vapor first enters the condensing chamber, it is suddenly cooled and is deposited as the fine powder known as flowers of sulphur. When the chamber becomes warmer, the vapor condenses in the form of a liquid, which is drawn off from the bottom of the chamber and is molded in wooden molds into the form of roll sulphur.

59. Properties. Sulphur is a yellow, brittle substance. It is insoluble in water, but will dissolve readily in the liquid known as carbon bisulphide. It behaves very peculiarly when heated. When the temperature reaches 114.5° C., the sulphur melts, forming a thin, straw-colored liquid. As the heat increases, the mass becomes darker in color, and at 200° to 250° C. it becomes so thick that the vessel in which it is heated can be turned upside down and the sulphur will not run out. Finally, it again becomes liquid, and at 448.4° C. it boils and is converted into a brownish yellow vapor.

60. Different Forms of Sulphur. If sulphur is dissolved in carbon bisulphide and the clear liquid is poured off and is allowed to evaporate slowly, crystals of sulphur (Fig. 68) will be formed, which, when examined through a magnifying glass, are seen to be *eight-sided*. This is the form of sulphur found in nature. If sulphur that has been gently heated to the melting point is allowed to cool until it is about half solidified, the solid part remaining after the liquid is poured off is found to be in the form of long *needlelike crystals* quite different from those described above (Fig. 69).


FIG. 68. — Rhombic or eight-sided sulphur crystals.

On the other hand, if sulphur is heated to the boiling point, and the liquid is poured in a thin stream into cold water (Fig. 70), it forms a plastic mass, entirely different in appearance from either of the other forms, and with no crystalline appearance. This is known as *plastic*, or *amorphous, sulphur*.


FIG. 69. — Prismatic or needlelike sulphur crystals.

Neither the needlelike (prismatic) sulphur nor the plastic sulphur is stable, because each, upon standing, gradually changes over into the eight-sided form. The greater stability of this eight-sided sulphur explains why it is the form found in nature.

Still another form of sulphur is known. This is found sometimes in sulphur springs, but is generally obtained by precipitating sulphur from some of its compounds. It is almost white in color and is known as *milk of sulphur* or *lac sulphur*. It is used in medicine.


FIG. 70. — The formation of amorphous sulphur.

It will thus be seen that sulphur can exist in several very different forms. Each of these forms, however, consists of sulphur and nothing else. These are known as *allotropic* forms of sulphur. The property of existing in different forms is known as *allotropy* from two Greek words meaning simply "another form."

61. Sulphur Found in Compounds. In addition to its occurrence in the elementary condition sulphur is found in many compounds. It occurs in the water of sulphur springs, in the air near volcanoes, and in many minerals. It is found in many plant and animal tissues. It is used as an insecticide and as a fungicide. It is an ingredient of gunpowder and of fireworks, and is used in the manufacture of hard and vulcanized rubber. Its most important use is in the manufacture of sulphuric acid, one of the most important of chemical substances.

62. Sulphides. Sulphur somewhat resembles oxygen in its chemical behavior, especially toward metals. Like oxygen it unites slowly with some metals at ordinary temperatures. Like oxygen, also, it readily combines with most metals at high temperatures, as can be shown in the following experiment. Four grams of flowers of sulphur and seven grams of fine iron filings are thoroughly mixed by rubbing in a mortar. That this is only a mechanical mixture may be shown by taking a portion of the mixture and drawing the iron out with a magnet, or by dissolving out the sulphur with carbon bisulphide. Another portion of the mixture is placed in an old test tube and heated over the Bunsen burner. As soon as the mass begins to glow, it is removed from the flame. If, when it is cold, the tube is broken and its contents are examined, it will be found that the magnet does not attract the iron in the mass, and that sulphur cannot be dissolved from it by carbon bisulphide. The sulphur has *combined chemically* with the iron, and an entirely new substance, known as *iron sulphide*, has been formed. Copper, lead, and other metals will combine with sulphur in the same way. The compound formed by the union of sulphur with another element is called a *sulphide*, just as that formed by the union of oxygen and another element is called an oxide. Sulphides are abundant in nature and many of them are *valuable ores*. Lead, copper, mercury, and zinc are among the metals found in nature as sulphides. The black substance which forms on a silver spoon that has been allowed to remain in contact with an egg is silver sulphide, the sulphur coming from the egg. Nearly all sulphides are insoluble in water.

63. Sulphur Dioxide. In the study of oxygen it was shown that sulphur burns, forming a suffocating gas. Even the sulphides of the metals will burn if heated, both the sulphur and the metal uniting with oxygen and yielding the gaseous oxide of sulphur just mentioned, and the oxide of the metal. Copper sulphide, for instance, when burned becomes copper oxide, the sulphur which was combined with the copper uniting with oxygen to form a gas. This gas has been named *sulphur dioxide* for reasons to be explained shortly.

Sulphur dioxide is a colorless gas with a suffocating odor. It is 2.21 times heavier than air. It can readily be condensed to a liquid. It will not burn nor support combustion. In the presence of moisture it bleaches many organic dyes. Flowers and moistened pieces of cloth colored with vegetable dyes placed in a bottle in which sulphur is burned will lose their color. Sulphur dioxide is used to bleach silk, wool, straw, and some other fibers which would be injured by some of the more powerful bleaching agents. It has been largely used as an *antiseptic* and *disinfectant,* as it will kill bacteria. A common method of disinfecting after illness is to burn sulphur in the closed room. The sulphur candle (Fig. 71) in which the brimstone is molded around a wick is a convenient form to use for this purpose. As a disinfectant it is now largely replaced by formaldehyde, however, which does not destroy the colors of the materials disinfected, nor injure the metal of the gas or electric fixtures, and other hardware. Sulphur dioxide has also been used in preserving fruits and other food products since it prevents fermentation; *but the use of this substance in anything that is to be eaten is objectionable.* Sulphur dioxide is injurious to plants, and it is not unusual to find trees and other vegetation completely destroyed in the vicinity of the smelting works where the sulphide ores are being roasted to burn off the sulphur.

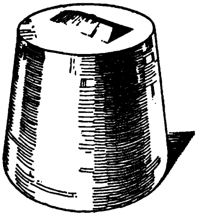


FIG. 71. — A sulphur candle.

64. Sulphur Trioxide. Sulphur dioxide has been found, upon analysis, to contain equal parts by weight of sulphur and oxygen. Under ordinary circumstances the sulphur dioxide does not readily take on any more oxygen. If, however, sulphur dioxide and oxygen are passed over finely divided platinum which is highly heated, a new compound of sulphur and oxygen will be formed which contains one part of sulphur to one and one half parts of oxygen, or one half more oxygen than is found in sulphur dioxide. This compound is a colorless liquid which solidifies at about 15° C., and has been named *sulphur trioxide.* It may be prepared on a small scale by means of the apparatus shown in Fig. 72. The finely divided platinum is prepared by moistening asbestos fiber with a solution of platinum chloride and igniting it in a flame. The platinum-asbestos is placed in the hard glass tubing at *A*. Air, to furnish oxygen, is introduced at *B*, and sulphur dioxide at *C*. The tube and asbestos are heated with a Bunsen burner, and the fumes of sulphur trioxide will be seen escaping into the air at *D*. It may be condensed to a liquid if conducted into a test tube surrounded with a freezing mixture of ice and salt.

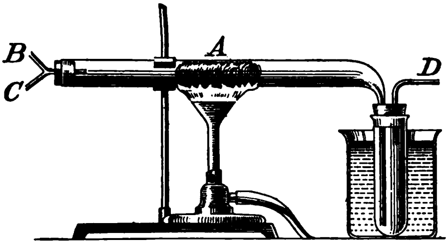


FIG. 72. — **Preparation of sulphur trioxide.**

It is to be noted that the platinum itself apparently does not take part in the chemical change, but in some unexplainable way causes the union of the oxygen and sulphur dioxide to take place very readily. Substances like this, which hasten what would otherwise be slow chemical changes, are called *catalytic agents* or *catalyzers,* and the action is called *catalysis.* A catalyzer, then, merely increases the speed of the chemical change but does not alter its products.

65. Sulphuric Acid. Sulphur trioxide unites very vigorously with water. When exposed to the air it fumes strongly, and if thrown upon water it hisses like hot iron. When sulphur trioxide unites with water, the product is *sulphuric acid,* sometimes called oil of vitriol. At the present time most of the sulphuric acid is manufactured by a method in which sulphur trioxide is first prepared by a process similar to the one described. Usually the sulphur dioxide is obtained by burning pyrites, which is a sulphide of iron. The sulphur dioxide and air are conducted into towers in which the formation of sulphur trioxide is brought about by means of platinum or some other catalyzer, and the sulphur trioxide is then united with water to form sulphuric acid. This is known as the *contact method of making sulphuric acid.* Another method will be mentioned in Chapter X.

Sulphuric acid is probably the most important of all manufactured chemicals. It is the most common reagent of the chemical laboratory; and is used in many industries, especially in the manufacture of fertilizers, soda, aniline dyes, and nitroglycerin, and in the refining of petroleum. Over four million tons of sulphuric acid are produced annually.

When sulphuric acid is pure it is a colorless, odorless, oily liquid. It is almost twice as heavy as water, having a specific gravity of 1.84. It unites with water with great evolution of heat. For this reason great caution must be observed in mixing sulphuric acid and water. The acid should be poured slowly into the water with constant stirring. The water should never be poured into the acid, for the great amount of heat liberated is likely to cause an explosion which will throw the acid out of the container. Sulphuric acid exposed to the air absorbs water vapor until it becomes quite dilute. It is, therefore, a good drying agent and is often used to dry gases. It destroys plant and animal tissue, for it *abstracts hydrogen and oxygen from them in proportions to form water,* leaving behind a black charred mass which is largely carbon. The painful burns caused by sulphuric acid are due to this action. It is also sometimes used as an oxidizing agent, for under certain conditions it will give up part of its oxygen to another substance.

In very dilute solutions sulphuric acid has a sour taste. If a piece of blue litmus paper, a paper colored with a vegetable dye called litmus, is placed in it, the paper will *turn red.* The sour taste and the effect on litmus are properties that are characteristic of a large group of chemical compounds which are known as *acids.*

66. Sulphuric Acid Forms Salts. Sulphuric acid contains sulphur, oxygen, and hydrogen, for it is produced by the combination of sulphur trioxide with water. In Chapter III, in the discussion of the preparation of hydrogen, it was shown that the hydrogen could be liberated from sulphuric acid. It will be well now to consider what became of the other elements in the acid.

It was noticed that while hydrogen was being evolved the zinc disappeared. If the liquid in the hydrogen generator is now filtered to remove the black particles which are due to impurities in the zinc, and is then slowly evaporated in a porcelain dish placed in a sand bath, the residue is a white solid which looks somewhat like common salt. If this material is redissolved in a small quantity of hot water and set aside for some time, clear crystals are formed. This substance contains the zinc which disappeared and the sulphur and oxygen which were in the sulphuric acid. It is a compound known as *zinc sulphate.* If iron is used instead of zinc, the residue is the green material sold under the name of copperas, which is in reality iron sulphate, and consists of iron, sulphur, and oxygen.

What has really happened in these cases, then, is that the zinc or iron has *replaced the hydrogen which was in the sulphuric acid.* A compound formed by replacing the hydrogen of sulphuric acid with a metal is called a *sulphate.* Some of the sulphates like zinc sulphate and iron sulphate can be made by the action of the acid on the metal, but other metals are not readily acted upon by sulphuric acid. The sulphates of nearly all metals, however, can be made by indirect methods; namely, *by the action of sulphuric acid on the oxide of the metal.* This property of sulphuric acid by which its hydrogen can be replaced by a metal is a property that is *common to all acids.* Many of the compounds so formed resemble common salt in appearance, and for that reason, it has become customary to call all such compounds *salts.* Copperas, blue vitriol, washing soda, Epsom salts, alum, and table salt are all common substances which the chemist places in the large class of compounds known as salts. *They are all formed by replacing the hydrogen of some acid with a metal.*

67. Multiple Proportions. It has been shown that sulphur forms two distinct compounds with oxygen. Each of these compounds is constant in composition in accord with the law of definite proportions (34). Hydrogen and oxygen likewise form two distinct compounds, water and hydrogen peroxide. In the latter one part of hydrogen is combined with sixteen parts of oxygen, while in water one part of hydrogen is combined with eight parts of oxygen. The ratio between the amounts of oxygen in these compounds when stated in its simplest form is as 1 to 2.

In sulphur dioxide one part of sulphur is united to one part of oxygen, while sulphur trioxide is composed of one part of sulphur to one and one half parts of oxygen. The ratio of the oxygen in these compounds when expressed in the simplest terms is as 2 to 3.

A study of all the cases where two elements unite in more than one proportion has shown that similar simple ratios always exist between the different amounts of one of the elements which unite with a fixed amount of the other; that is, the ratio can always be expressed by small whole numbers such as 1, 2, 3, 4, or 5. This fact is known as *the law of multiple proportions.*

EXERCISES

Ex. 47. Examine a piece of roll sulphur and state its most obvious physical properties. Will it dissolve in water? In carbon bisulphide? Heat the sulphur in a test tube. What happens upon heating? Is sulphur an element or a compound? How does it occur in nature? How is native sulphur purified?

Ex. 48. Dissolve one half gram of sulphur in 2 cubic centimeters of carbon bisulphide. Pour the liquid into a small dish and allow the carbon bisulphide to evaporate. Describe the crystals of sulphur which form.

Ex. 49. Fold a piece of filter paper as if to be placed in a funnel. Heat about half a test tube of sulphur until it just melts and, holding the filter by the three folds, pour the melted sulphur into it. When the crystals begin to form across the surface pour off the remaining liquid. What is the form of the crystal in this case? Allow the crystals to remain undisturbed for a few days and examine again. Has any change taken place?

Ex. 50. Heat another portion of sulphur to the boiling point and pour in a thin stream into cold water, moving the test tube so as to form a coil rather than a solid mass. Describe the result. Is this sulphur crystalline? Define amorphous. Preserve the sample and note what change takes place in a day or two. Is this form of sulphur stable? What is the stable form of sulphur? What is meant by allotropic forms of sulphur? What is the property of an element by which it can exist in different forms?

Ex. 51. Rub together four grams of powdered sulphur and seven grams of fine iron filings. Pass a magnet over a small portion of the mixture. Treat another small portion with carbon bisulphide, pour off the liquid, and evaporate. Have you a compound of iron and sulphur or a mechanical mixture? Place another portion in an old test tube and heat over the Bunsen burner. What happens? When cold, break the tube and test the contents with the magnet and with carbon bisulphide. What results do you get? Has a chemical change taken place? What is the new compound called? Are compounds of sulphur of frequent occurrence? Name some substances containing sulphur. How does sulphur resemble oxygen in chemical behavior? What are the compounds of sulphur with the metals called? What makes the silver spoon turn black when used with eggs? Why does a silver coin become black when kept in the pocket with a rubber band? Silver jewelry sometimes turns black; where does the sulphur come from?

Ex. 52. Burn a small quantity of sulphur in a bottle of air (or oxygen if at hand). Smell the gas cautiously. What is this gas called and of what is it composed? Heat a piece of iron sulphide, lead sulphide, or copper sulphide in the Bunsen flame. What odor do you note? What is the source of the sulphur? Give the physical properties of sulphur dioxide. Place some flowers and some pieces of moist colored cloth in the bottle of sulphur dioxide. What happens? What use is made of this property of sulphur dioxide? How is sulphur dioxide used as a disinfectant? Should sulphur dioxide be used to preserve fruits and vegetables? What effect does sulphur dioxide have upon vegetation?

Ex. 53. (By the Teacher.) Perform the experiment described in paragraph 64. If a cylinder of commercial sulphur dioxide is not at hand, the gas may be generated by the action of sulphuric acid on sodium sulphite. (Convince the class that this gas is the same as that obtained by burning sulphur.) What is the product obtained in this experiment? How does the proportion of oxygen in this compound compare with that found in sulphur dioxide? What are the properties of sulphur trioxide? Why was the platinum used in this experiment? What is meant by a catalyzer? What other example have you had of a catalytic agent?

Ex. **54.** What compound is formed when sulphur trioxide unites with water? Examine the sulphuric acid of the laboratory and state the physical properties. Pour ten cubic centimeters of sulphuric acid slowly into twice as much water. What happens? Should the water ever be poured into the acid? Why? What happens when sulphuric acid stands exposed to the air? If moist air or other gas were forced through sulphuric acid, would it come out drier or more moist than before? Put a few drops of the acid on a clean piece of wood. What happens? Dip a bit of cotton cloth, a leaf of a plant, and a feather in strong sulphuric acid and describe the result. What is the cause of these changes? What would be the effect of spilling sulphuric acid on your flesh? On your clothing? Tell something of the importance of sulphuric acid. What is a common name for it? What is meant by the contact method of making this acid?

Ex. **55.** Put eight or ten drops of sulphuric acid in a tumblerful of water. Taste it cautiously by dipping in a glass rod and touching it to the tongue. How does the mixture taste? Dip a piece of blue litmus paper into the liquid. What change takes place? Of what group of compounds are the sour taste and this effect on litmus paper characteristic?

Ex. **56.** Perform the experiment described in paragraph 66. What is the appearance of the residue left after evaporation? Of what is the material composed? What is it called? What would be formed if iron were used in place of zinc? Give a general definition for sulphates. What does the chemist mean by a "salt"? How are all salts formed?

CHAPTER VIII

THE ATOMIC THEORY

68. The modern science of chemistry was preceded by the work of the alchemists during the Middle Ages. These men were striving to find a method by which the common metals could be converted into gold, and while they failed in the particular thing they desired, they discovered many substances which have been of great value to mankind. Sulphuric acid, nitric acid, and hydrochloric acid, as well as certain of the methods of extracting metals from their ores, are discoveries which were due to the work of the alchemists. But the methods of these men were haphazard, and it was not until the eighteenth century, when the use of the balance made quantitative studies possible, that the science now known as chemistry had its beginning. It was during this period that the law of definite proportions (34) and the law of multiple proportions (67) were discovered. These laws, it should be understood, are merely *concise statements of truths that have been proved by experiment.*

It is one thing, however, to know a general fact, and quite another thing to know the cause of the fact. While it is known that elements combine in definite and multiple proportions, it does not necessarily follow that it is known why they combine according to these laws. It is natural for man to desire to know the reason for the truths which he discovers, and when the cause cannot be determined directly he *imagines a cause,* or a condition which, if it existed, would lead to the results discovered. Such a theoretical explanation is known as a *hypothesis.*

If, now, this hypothesis is tested in every way that suggests itself and all facts discovered are in accordance with it, it becomes a *theory.* A hypothesis is a guess in regard to the cause of certain phenomena, while a theory is a hypothesis which has been thoroughly tested and found to be fully in accord with the known facts.

69. The Atomic Theory. The explanation of the laws of definite and multiple proportions, now generally accepted, is the hypothesis formulated by Dalton, the English chemist, about 1804, which has come to be known as the *atomic theory.* This theory assumes, (1) that all elements are made up of minute independent particles called *atoms* that cannot be subdivided, (2) that all atoms of the same element have the *same size and weight,* while the atoms of different elements have different weights, (3) that when two or more elements unite, the action consists in the union of a definite small number of the atoms of each element to form a small particle of the compound.

70. Molecules. The atom is the smallest particle of an element, but it is evident that the smallest particle of a compound which can exist must contain more than one atom; that is, it must contain at least one atom of each element in the compound and may contain more than one atom of each element. This smallest particle of a compound, which consists of two or more atoms, is called a *molecule.* Even in the case of the elements it is not possible for the atoms to exist in the free state, but these also combine into groups or molecules. There are, therefore, two kinds of molecules; namely, the molecules of an element in which all the atoms are alike, and the molecules of a compound in which there are at least two different kinds of atoms. The molecule of hydrogen is supposed to contain two atoms of hydrogen, and the molecule of water to contain two atoms of hydrogen and one of oxygen.

71. The Atomic Theory Applied. If the atomic theory is accepted, it is easy to understand why elements unite according to the laws of definite and multiple proportion. It is assumed that the molecule of water contains one atom of oxygen. If any more oxygen is added to the molecule, it must be at least one whole atom, for *the atom is indivisible;* but the addition of one more atom would exactly double the amount of oxygen in the molecule. The facts are in accord with the theory, for hydrogen peroxide contains exactly double the quantity of oxygen that is present in water. Sulphur dioxide is assumed to contain one atom of sulphur and two atoms of oxygen, and sulphur trioxide to contain one atom of sulphur and three atoms of oxygen. This assumption is in accord with the fact that the ratio of oxygen in the two compounds is as 2 to 3. It must not be forgotten, however, that the laws of definite and multiple proportions are *truths,* but that the atomic theory, although it is a conception which is probably true, cannot be proved to be true. All the facts of chemistry discovered since the time of Dalton, however, are in perfect accord with the theory, and by means of this theory chemists were able to predict many of the facts which have since been discovered. At the present time the existence of molecules and atoms can scarcely be doubted.

72. Atomic Weights. Atoms are too small to be weighed directly, but if the atoms of each element are all alike and of the same weight, it ought to be possible to assign some number to the elements which would represent the *relative weights of the atoms.* This was first done by giving to hydrogen, which was the lightest atom, the relative weight of one, and studying the combinations which the other elements made with it, or with some other element whose relation to hydrogen had already been established. Most of the atomic weights have been determined by studying the combinations of the various elements with oxygen, which had been determined to be 16; that is, the atom of oxygen was said to be 16 times as heavy as the atom of hydrogen. As a matter of fact the atomic weights are now all based on the assumption that oxygen is 16, for it has been found that there was a slight error in the earlier calculations and that hydrogen is 1.008 if oxygen is 16.

73. Chemical Symbols. In order to avoid the inconvenience of using long names the chemist represents the different elements by symbols. These symbols consist of the first letter of the name, unless there is more than one element with the same initial letter. In that case the first element discovered is designated by the initial letter, and the others by the initial letter with some other characteristic letter of the name; as, for instance, C is the symbol for carbon, but Cl for chlorine and Ca for calcium. Chemical symbols are the same the world over, hence where the initial letter of the name differs in different languages an abbreviation of the old Latin name is used for the symbol. Fe (ferrum) is the symbol for iron, and Cu (cuprum) for copper.

The symbol designates not merely the element, but stands in each case for one atom of the element. Thus, H represents one atom of hydrogen, O one atom of oxygen, S one atom of sulphur. But the symbols mean even more than this, for they express the atomic weights as well. Thus, O not only means one atom of oxygen, but also means that this atom weighs sixteen times more than one atom of hydrogen. The table on the inside of the back cover gives a list of the elements, with their symbols and atomic weights. A few of the more common elements, with their atomic weights stated in round numbers, are given in the following table:

ELEMENT	SYMBOL	ATOMIC WEIGHT
Hydrogen	H	1.
Oxygen	O	16.
Nitrogen	N	14.
Sulphur	S	32.
Carbon	C	12.
Calcium	Ca	40.
Iron	Fe (Ferrum)	56.
Sodium	Na (Natrium)	23.
Chlorine	Cl	35.5
Phosphorus	P	31.
Potassium	K (Kalium)	39.
Copper	Cu (Cuprum)	63.6
Zinc	Zn	65.4

If more than one atom is to be designated the proper numeral is placed before the symbol, thus:

2 H means 2 atoms of hydrogen.

3 S means 3 atoms of sulphur.

But if the atoms are combined with others in a compound a small numeral is placed after and below the symbol, thus:

H_2 means 2 atoms of hydrogen in combination.

S_3 means 3 atoms of sulphur in combination.

EXERCISES

Ex. **57.** State the law of the conservation of matter. Is the amount of matter in the universe constant? Does any matter disappear when coal burns in the stove? Is the amount of each chemical element in the universe always constant? What is the law of definite proportions? Of multiple proportions? What is meant by a hypothesis? A theory?

Ex. **58.** State the atomic theory. Show how the atomic theory explains the law of multiple proportions. What name is given to the smallest amount of a compound which can exist? Are there molecules of the elements as well as of the compounds? What is meant by the atomic weight of an element? By a chemical symbol? Of what does the chemical symbol consist? Just what does the symbol for an element, S for example, designate?

CHAPTER IX

FORMULAS AND EQUATIONS

74. Chemical Formulas. A formula is a group of symbols that is used to express the *composition of a molecule of a compound.* The symbols of the different elements are written side by side, with the proper subscript numerals representing the number of atoms of each element. Thus, H_2O is the formula for water, and indicates that a molecule of water is composed of two atoms of hydrogen and one of oxygen. It also indicates that water is composed of 2 parts by weight of hydrogen and 16 parts by weight of oxygen. Sulphur dioxide is written SO_2, and the trioxide SO_3. As the atomic weight of sulphur is 32, it is seen that the dioxide consists of 32 parts by weight of sulphur to 32 of oxygen. The ratio in the trioxide is 32 of sulphur to 48 of oxygen. The formula also shows why one is called *di*-oxide and the other *tri*-oxide. The formula for sulphuric acid is H_2SO_4, showing that the molecule contains 2 atoms of hydrogen, 1 of sulphur, and 4 of oxygen, the relative weights being as 2 to 32 to 64. The sum of the atomic weights in a molecule is the *molecular weight.* The molecular weight of sulphuric acid is 98.

A numeral placed in front of a formula multiplies the molecule and consequently all the atoms within the molecule. Thus, $3\,H_2SO_4$ represents three molecules of sulphuric acid and is equivalent to 6 atoms of hydrogen, 3 of sulphur, and 12 of oxygen. In certain cases a group of elements which act together are inclosed in a parenthesis and the group may be multiplied by using the subscript to the right of the parenthesis; thus, in $Fe_2(SO_4)_3$ the group SO_4 is multiplied by 3. If a numeral is placed before this formula, as $2\,Fe_2(SO_4)_3$, it means that the whole formula is multiplied by 2, and that consequently there are six SO_4 groups present and four atoms of iron.

75. Reactions. A chemical reaction is a special or limited chemical change. When sulphuric acid acts on zinc to produce hydrogen the change which takes place is a chemical reaction. The burning of sulphur to produce SO_2, the uniting of oxygen and copper to form copper oxide, the decomposition of water by means of heated iron, are all chemical reactions.

76. Reagents. A reagent is a substance capable of producing a reaction with another substance. The name is sometimes confined to those chemicals which are employed to detect the presence of other substances.

77. Chemical Equations. To express the various facts about chemical reactions it is customary to use a sort of *chemical shorthand* known as a *chemical equation.* The formulas representing the substances which enter into the reaction are connected by the plus (+) sign and form the left-hand member of the equation. The formulas for the products of the reaction are placed at the right, and the equation is read from left to right. In place of the sign of equality as used in mathematical equations it is now customary to use the arrow to connect the two members of the equation. For example, it is known that when water vapor is passed over heated zinc the result is the production of zinc oxide and hydrogen (21). This fact may be represented by the following equation:

$$H_2O + Zn \rightarrow ZnO + 2\,H.$$

The equation should not be read like an equation in mathematics, nor should it be considered as any more than a brief way of stating certain known facts. The plus sign should be read " and " and the arrow translated as " yields " or " produces." The above equation, then, means that *water and zinc act upon each other to produce zinc oxide and hydrogen.* It also shows that one molecule of water reacts with one atom of zinc and produces one molecule of zinc oxide and two atoms of hydrogen.

78. Writing Reactions. If it is remembered that the equation is merely a brief way of expressing certain facts, it will be understood that *a complete knowledge of the reaction is necessary before the equation can be written.* The equation cannot be figured out mathematically, nor can the equation for one element be " guessed out " by knowing what another will do under the same circumstances. When copper burns in oxygen, the equation is

$$Cu + O \rightarrow CuO.$$

But when iron burns the equation is as follows:

$$3\,Fe + 4\,O \rightarrow Fe_3O_4.$$

In the case of phosphorus it is

$$2\,P + 5\,O \rightarrow P_2O_5.$$

The reaction for the formation of water may be expressed as follows:

$$2\,H + O \rightarrow H_2O.$$

The preparation of oxygen is represented by the following equation:

$$KClO_3 \rightarrow KCl + 3\,O.$$

This equation should be read: *Potassium chlorate yields potassium chloride and oxygen.* The manganese dioxide is not written into the equation because it remains unchanged and probably acts as a catalyzer. The following equations will express the reactions mentioned in the last chapter.

Formation of sulphides:

$$3\ \mathrm{Fe} + 4\ \mathrm{S} \rightarrow \mathrm{Fe_3S_4};$$
$$\mathrm{Cu} + \mathrm{S} \rightarrow \mathrm{CuS}.$$

Burning of sulphur or a sulphide:

$$\mathrm{S} + 2\ \mathrm{O} \rightarrow \mathrm{SO_2};$$
$$\mathrm{CuS} + 3\ \mathrm{O} \rightarrow \mathrm{CuO} + \mathrm{SO_2}.$$

When sulphur dioxide changes to trioxide:

$$\mathrm{SO_2} + \mathrm{O} \rightarrow \mathrm{SO_3}.$$

The platinum is not written into the equation because it acts as a catalyzer and takes no actual part in the reaction. When sulphuric acid is formed, the equation is

$$\mathrm{SO_3} + \mathrm{H_2O} \rightarrow \mathrm{H_2SO_4},$$

and when sulphuric acid and zinc are used to prepare hydrogen,

$$\mathrm{Zn} + \mathrm{H_2SO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{H_2}.$$

This equation should be read: *zinc and sulphuric acid yield zinc sulphate and hydrogen.*

Since the atoms are indestructible, it follows that the same number of atoms of each element should be found on both sides of the arrow. The equations are also *quantitative* expressions. The last equation expresses the fact that 65.4 parts by weight of zinc will react with 98 parts of sulphuric acid and produce 161.4 parts of zinc sulphate and 2 parts of hydrogen.

Thus if the quantity of any one of the factors of the equation is known, it will be seen that any or all of the others can be calculated. For example, suppose the problem is to find how much sulphuric acid is required to produce 7 pounds of hydrogen. First write the equation with the molecular weights (73) written below:

$$\begin{array}{ccccccc} Zn & + & H_2SO_4 & \rightarrow & ZnSO_4 & + & H_2, \\ 65.4 & & 98 & & 161.4 & & 2. \end{array}$$

The equation shows that 2 parts of hydrogen can be produced from 98 parts of sulphuric acid; this, therefore, establishes the ratio, and the problem is merely a matter of simple proportion:

$$2:98 = 7:x, \qquad x = 343.$$

Consequently 343 pounds of sulphuric acid are required to produce the 7 pounds of hydrogen. To solve similar problems, first write the equation with the correct atomic or molecular weights, and then state the problem in the form of a proportion, like the one given above.

79. Chemical Affinity. It is not known why certain substances act upon each other chemically and others do not. The fact that a piece of sulphur will burn when heated and platinum will not is well known, but why this is so no one can tell. For want of a better name this force, or attraction, is called *chemical affinity.* Whatever this force is that holds the elements together, it is very important; for without it the compounds could not exist, and if it ceased to act all the complex substances of the animal, vegetable, and mineral kingdoms would dissociate into a few simple substances known as elements. Elements that readily unite are said to have great affinity for each other. Sulphur, since it unites with oxygen, is said to have affinity for oxygen.

Platinum, which cannot be made to burn in air, is said to have slight affinity for oxygen.

80. Valence. The power that an atom of one element has to unite with one or more atoms of another element is called its *valence.* Here, again, hydrogen is used as the standard and is rated at 1. Any atom that can hold one atom of hydrogen in combination is said to have a valence of 1, or to be *univalent.* If it can combine with 2 atoms of hydrogen, it is *bivalent.* Oxygen is bivalent because it unites with 2 atoms of hydrogen (H_2O). Elements which do not unite with hydrogen are compared with oxygen or some other element which does unite with hydrogen. An atom of copper unites with one atom of oxygen and consequently must be bivalent. The zinc atom does not unite with hydrogen, but since it replaces the 2 atoms of hydrogen in sulphuric acid, it is bivalent. The matter of valence would be very simple if all elements had just one valence, but some of them vary in valence. Sulphur is apparently quadrivalent in SO_2 and hexavalent in SO_3. When an element exhibits more than one valence, it is generally true that *the compounds at one of the valences are much more stable.* The compounds in which sulphur has a valence of six are much more stable than those in which it appears to have a valence of four.

81. Hydrogen Peroxide. The stable combination of hydrogen and oxygen is water, H_2O, sometimes written H—O—H. In hydrogen peroxide another atom of oxygen is introduced into the molecule, making it H—O—O—H or H_2O_2. This oxide of hydrogen readily changes to the more stable compound water, and one atom of oxygen is liberated. For this reason, hydrogen peroxide, or more properly *dioxide,* is a *strong oxidizing agent.* It gradually decomposes into water and oxygen upon standing, but does so very quickly if in contact with a substance that can be oxidized. The equation is

$$H_2O_2 \rightarrow H_2O + O.$$

82. Physics and Chemistry. From what has been said in this chapter it must be evident that so long as the molecule remains intact the chemical composition of a substance does not change. If the make-up of the molecule changes, new substances are formed, and the change which takes place is a *chemical change*. It may be said, then, that physical changes are those in which the composition of the molecule is not affected, while in chemical changes the atoms are rearranged into new and different molecules.

EXERCISES

Ex. 59. What is a chemical formula and what does it represent? What is the formula for sulphuric acid? What facts about sulphuric acid are represented by this formula? What is meant by the molecular weight of a compound? The formula for potassium chlorate is $KClO_3$; what is its molecular weight? What percentage of oxygen does it contain?

Ex. 60. What is meant by a chemical equation? Of what use are these equations? How should they be read? What must be known in order to write an equation? How much potassium chlorate would be needed to produce 100 pounds of oxygen?

What is meant by chemical affinity? By valence? The formula for carbon dioxide is CO_2; what is the valence of carbon? What is the relation of the molecule to chemical and physical changes?

CHAPTER X

ACIDS OF SULPHUR AND HYDROGEN SULPHIDE

83. Sulphurous Acid. Sulphur dioxide dissolves readily in water, one volume of the latter absorbing forty volumes of the gas. This liquid *has a sour taste and turns blue litmus paper red,* from which it may be inferred that the solution contains an acid. The reaction may be expressed as follows:

$$SO_2 + H_2O \rightarrow H_2SO_3.$$

The compound H_2SO_3 is *sulphurous acid.* When the solution is heated, the acid decomposes into water and sulphur dioxide, the latter being driven off. This reaction may be written:

$$H_2SO_3 \rightarrow H_2O + SO_2.$$

The fact that sulphurous acid decomposes when heated makes it impossible to concentrate it and obtain it free from water, as can be done with sulphuric acid, which does not completely decompose upon boiling. Sulphurous acid possesses marked bleaching properties. In fact the bleaching and disinfecting properties which were referred to in describing sulphur dioxide should properly be ascribed to sulphurous acid, for it was noted that these effects were produced when the substance acted upon was moist. Under these circumstances the sulphur dioxide unites with the moisture to form sulphurous acid.

Sulphurous acid can readily be oxidized to sulphuric acid:

$$H_2SO_3 + O \rightarrow H_2SO_4.$$

This change is brought about to a limited extent by the oxygen of the air uniting with the sulphurous acid. The change can also be effected by means of oxidizing agents, as, for instance, hydrogen peroxide:

$$H_2SO_3 + H_2O_2 \rightarrow H_2SO_4 + H_2O.$$

This property of readily taking on oxygen makes sulphurous acid *a strong reducing agent.*

84. Sulphuric Acid by Chamber Process. The older method of preparing sulphuric acid, known as the *chamber process,* depended upon the oxidation of sulphurous acid by means of one of the oxides of nitrogen. Under certain circumstances this compound will give up a part of its oxygen, and later will take up again a like quantity of oxygen from the air. This oxide may be represented by the formula NO_2. In manufacturing sulphuric acid, sulphur or pyrites is burned to produce SO_2, and this compound and steam are conducted into lead-lined chambers. The lead lining is used because sulphuric acid has very little effect on lead. The sulphur dioxide and steam unite to form sulphurous acid. The NO_2 gas which is obtained from nitric acid then acts upon the sulphurous acid:

$$H_2SO_3 + NO_2 \rightarrow H_2SO_4 + NO.$$

Air is also admitted, and the gas takes up oxygen and changes back to NO_2:

$$NO + O \rightarrow NO_2.$$

Theoretically the same amount of NO_2 could be used indefinitely, as it acts much like a catalyzer. Practically there is always some loss of the nitrogen oxide, since some of it is dissolved in the sulphuric acid and removed from the chamber with it. The actual chemical changes which take place in the chamber are much more complicated than are indicated in the foregoing equations, and are not fully understood, but the essential feature of the process is this power of the oxide of nitrogen to take an intermediate part in the reaction between sulphurous acid and the oxygen of the air, or, as it is often expressed, to act as a *carrier of oxygen.* This process results in sulphuric acid of 50 to 60 per cent strength and with some impurities, namely, lead and nitrogen compounds. Where pure acid is required the contact method is generally used. The chamber method is used by many of the fertilizer factories as the so-called chamber acid is about the strength needed in the making of fertilizers, and the little impurity found in the acid is of no moment. Fully one half of all the sulphuric acid produced in the world is used in the manufacture of fertilizers.

85. Two Acids of Sulphur. It will be noticed that there are two acids containing hydrogen, sulphur, and oxygen, the difference in the formulas being one atom of oxygen. These acids are H_2SO_3 and H_2SO_4. Some other elements also form more than one acid in which the molecules differ only in the amount of oxygen present. To distinguish between these acids it is customary to use the ending "*ous*" for the acid having the smaller amount of oxygen, and the ending "*ic*" for the acid with the larger amount, hence, *sulphurous acid* for H_2SO_3 and *sulphuric acid* for H_2SO_4. In studying sulphuric acid, it was found that the hydrogen of the acid could be replaced by a metal to form a saltlike substance. The same thing holds true of sulphurous acid in a more limited way. The best-known compound of this class is sodium sulphite, which is commonly used in photography. It has the formula Na_2SO_3. The sodium atom is univalent, hence two atoms are required to replace the two hydrogen atoms of sulphurous acid. The sulphites, like sulphurous acid, are strong reducing agents. They take on oxygen and change to sulphates, thus:

$$Na_2SO_3 + O \rightarrow Na_2SO_4.$$

This change takes place slowly when the salt is exposed to the air, and consequently nearly all samples of sulphites contain some sulphates unless carefully protected. The salt of an acid ending in " ous " has the suffix *ite;* the salt of an acid ending in " ic " has the suffix *ate;* hence, the terms sulphite and sulphate.

86. Sulphuric Acid Used to Prepare Other Acids. When sulphuric acid is added to the salts of other acids, the metal of the salt and the hydrogen of the sulphuric acid change places, and a sulphate and the acid of the original salt are formed. If sulphuric acid is added to sodium sulphite, this reaction takes place:

$$Na_2SO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2SO_3.$$

As sulphurous acid happens to be a very unstable acid and is readily decomposed into sulphur dioxide and water, the equation is generally written to express that fact; thus,

$$Na_2SO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + SO_2.$$

The method used for the preparation of sulphurous acid is of general application, and many of the well known acids are prepared by *the action of sulphuric acid on the salt of the desired acid.* This method will be used repeatedly in the laboratory.

87. Hydrogen Sulphide. If dilute sulphuric acid is added to a sulphide, like one of the sulphides of iron, a gas is given off which is colorless, but which has a very offensive odor suggestive of rotten eggs. This gas, which has the formula H_2S, has been named *hydrogen sulphide* although it is popularly known as *sulphureted hydrogen.* The following equation represents its formation:

$$FeS + H_2SO_4 \rightarrow FeSO_4 + H_2S.$$

This gas exists naturally in sulphur springs, and in the air in the vicinity of volcanoes. It is poisonous and even in small quantities causes headache and nausea. It is slightly soluble in water, and the solution is frequently used in the analytical laboratory. This solution *turns blue litmus paper red* and in other ways gives evidence of being an acid; it is, therefore, often called *hydrosulphuric acid.* Here, then, is an acid which contains no oxygen, from which fact it may be inferred that Lavoisier was mistaken in his idea that oxygen is found in all acids. It will be seen later that there are other acids which contain no oxygen, one of them, hydrochloric acid, being commercially very important. Hydrogen sulphide is often formed during the decay of organic matter containing sulphur, especially of eggs and other animal matter. As most metallic sulphides are insoluble, hydrogen sulphide added to a solution of a salt of the metal will cause a precipitation of the insoluble sulphide. When H_2S is added to copper sulphate, the following reaction takes place:

$$CuSO_4 + H_2S \rightarrow CuS + H_2SO_4.$$

The copper sulphide, being insoluble, is precipitated. A piece of paper which has been moistened with a solution of a lead salt, such as lead acetate (sugar of lead), turns black if exposed to the hydrogen sulphide, owing to the formation of lead sulphide. *This is a test for hydrogen sulphide.*

88. Chemical Tests. A chemical test is a reaction used to recognize or detect the presence of a particular element or compound. The blackening of lead acetate paper shows the presence of hydrogen sulphide. The test for hydrogen is the fact that it burns and forms water; for oxygen, that it causes a glowing splint to burst into flame. There is no simple test for nitrogen because it is so inactive. The test for sulphur is the fact that SO_2 is formed upon burning; for a sulphite, that SO_2 is given off when a strong acid is added to it; for sulphuric acid or a sulphate, that a white precipitate, which will not dissolve in hydrochloric acid, is formed when a solution of barium chloride is added. Other tests will be discussed later.

EXERCISES

Ex. **61.** Place about two ounces of water in a wide-mouth bottle and shake the bottle so as to moisten its sides. By means of the crayon cup (Fig. 56) burn a small quantity of sulphur in the bottle, keeping the mouth covered. When the sulphur has stopped burning, remove the crayon cup and shake the bottle vigorously. Have you any evidence that the sulphur dioxide has dissolved in the water? Taste a drop of the liquid. Test it with blue litmus paper. What are the results? Does the liquid contain an acid? Write the equation for the formation of the acid. Heat a few drops of the liquid in a test tube. Is SO_2 given off? Write the reaction. Can sulphurous acid be prepared in concentrated form? Why? Dip a piece of colored calico in the liquid and note the effect. Will dry SO_2 bleach? What conditions are necessary to bleach with burning sulphur?

Ex. **62.** Add a few drops of sulphuric acid to a little water in a test tube. Now add a few drops of the laboratory solution of barium chloride. A white precipitate which will not dissolve in hydrochloric acid forms immediately. This is a test for sulphuric acid. To a little of the solution of sulphurous acid add the barium chloride solution. No precipitate will form. To another portion of sulphurous acid add a few drops of hydrogen peroxide and then add barium chloride. Have you any proof that sulphuric acid has been formed? Write the reaction. What is the principle upon which the chamber method of making sulphuric acid depends? What is the purpose of the oxide of nitrogen? Why is it called a carrier of oxygen? Is the acid made by the chamber process pure? What is the strength of chamber acid? Is much sulphuric acid used in the manufacture of fertilizers?

Ex. **63**. Give the formulas for the two acids of sulphur. When an element forms two acids, how are they named? What are the salts of sulphurous acid called? What happens to sodium sulphite upon standing exposed to the air? When sulphuric acid is added to sodium sulphite, what change takes place? Write the reaction. Is this a general method for preparation of acids?

Ex. **64**. Arrange apparatus as in Fig. 73. Place in the bottle a few pieces of iron sulphide. Through the thistle tube add dilute sulphuric acid. Describe the gas which is evolved. Write the reaction. Place the end of the delivery tube in a bottle half full of water and allow the gas to bubble through for some time. Test the solution with litmus paper. Have you any proof that an acid is present? What is it called? Does it contain any oxygen? How is hydrogen sulphide formed in nature? Where is it found naturally? To a solution of copper sulphate (blue vitriol), and to a solution of lead acetate, add a little of the hydrogen sulphide solution. What happens? Are most of the sulphides soluble or insoluble?

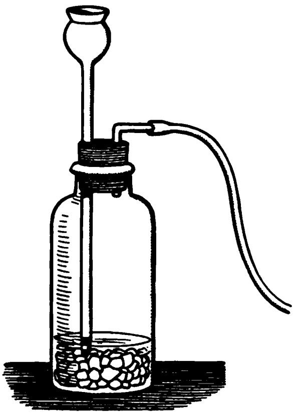


FIG. 73. — **Apparatus for the production of hydrogen sulphide.**

Ex. **65**. Explain how you could test for hydrogen sulphide. How could you tell whether a solid was a sulphide? Try the test with lead acetate paper. What test is used to detect sulphur? A sulphite? Sulphuric acid and sulphates? How can you tell whether a gas is hydrogen or oxygen?

CHAPTER XI

CARBON

CARBON is an element known to everyone in the form of the diamond, graphite or black lead, charcoal, lampblack, and coal. Like sulphur it exists in several allotropic forms.

89. Diamond and Graphite. The *diamond* is a brittle crystalline form of pure carbon. It is insoluble in all liquids and is the hardest substance known. If it is heated to a very high temperature in such a way that air is excluded, it swells and is converted into a black mass. Heated to a high temperature in oxygen, it burns completely, *yielding only carbon dioxide.*

Graphite, also called black lead, and plumbago, is a soft, shiny, black solid which is smooth and soapy to the touch. Pure graphite contains nothing but carbon. It exists in crystals, but the crystalline form is different from that of the diamond. Like the diamond, *it produces only carbon dioxide* when burned. It is used in the manufacture of lead pencils and stove polish, and as a lubricant where oil cannot well be used.

90. Wood Charcoal. The diamond and graphite are the only pure crystalline forms of carbon, but the element is known in many impure and amorphous forms. One of the best known of these is ordinary wood charcoal. Wood consists largely of carbon united with oxygen and hydrogen. If it is heated without access of air, the oxygen and hydrogen and part of the carbon are driven off in various liquid and gaseous compounds, and the remainder of the carbon is left behind in the form of charcoal. It is almost pure carbon, the only impurity being the small amount of mineral matter which it contains. Charcoal is a black, brittle substance, that often retains the form of the wood from which it was made. It is insoluble in water or acids, and burns without flame or smoke. It resists the action of chemicals and of decay bacteria; hence fence posts, and telegraph and telephone poles are often charred before being put into the ground.

Wood charcoal is now commonly made by heating wood in closed iron retorts, no air whatever being admitted. By this method the volatile products can be condensed and saved. Among the volatile substances of value produced in this way are *wood alcohol* and *acetic acid.* Such a process as this is known as *destructive distillation.*

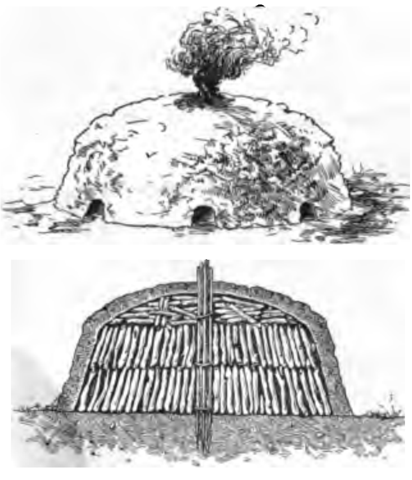


FIG. 74. — The exterior and section of a charcoal furnace.

The older method of making charcoal, and the one still largely used, is to construct a large pile of wood (Fig. 74) so arranged as to leave spaces between the pieces. The pile is covered with sods and earth to prevent free access of air, although small holes are left at the bottom and a large one at the top. The wood is lighted at the bottom and the fire is so controlled that it will smolder. The burning of the wood at the bottom of the pile heats the wood above sufficiently to drive off the volatile matter. After some time the holes are all closed to smother the fire, and then the pile is uncovered and the charcoal is removed. A very pure form of charcoal for use in medicine is sometimes produced from white sugar.

91. Animal charcoal is most commonly made by heating bones in closed retorts. This form is also known as *bone black*. Unless treated with acid to dissolve the mineral matter of the bone it contains only about 10 per cent of carbon. Animal charcoal is produced also from dried blood, a process by which a much purer form of carbon is obtained.

All these forms of charcoal have the power of absorbing offensive gases and are, therefore, used as deodorizers. Charcoal filters are used to remove objectionable substances from water. Charcoal also removes certain coloring matters from solutions. This property of charcoal (especially of animal charcoal) is utilized in refining sugar. The colored solution obtained from beet or sugar cane is passed through bone-black filters which remove the color, making possible the production of a white sugar. Charcoal is used also in the manufacture of gunpowder (207).

92. Coal. The different varieties of coal, which were formed by the gradual decomposition of vegetable matter in an insufficient supply of air, are forms of amorphous carbon. The vegetable origin is often shown by the fossil remains of leaves and stems of plants found in the coal. All forms of coal contain other substances in addition to the carbon. The different varieties of coal are commonly classified as hard and soft coals. Hard coal, or anthracite, is hard and lustrous. It is ignited with difficulty and burns with little flame, producing an intense heat. It contains about 95 per cent of carbon. Soft, or bituminous, coal burns with a smoky flame, and much volatile matter is produced, as can be seen by watching the little jets of flame which dart from a piece of burning soft coal. It contains about 80 per cent of carbon and, therefore, has a larger percentage of other compounds than has hard coal.

93. Coke is made by heating soft coal in an air-tight apparatus. Large quantities of coke are produced in the manufacture of illuminating gas. The bituminous coal is placed in large retorts and heated until all volatile matter is driven off, the material remaining in the retort being coke. It will thus be seen that coke bears the same relation to coal that charcoal does to wood. In addition to gas, referred to above, the heating of soft coal drives off coal tar, ammonia, and other volatile substances, which are utilized in modern coking plants. Many thousand tons of coke are made in the so-called " beehive " ovens (Fig. 75), where no attempt is made to save the volatile products. Millions of dollars worth of valuable products are wasted annually in this way.

FIG. 75. — Beehive coke ovens.

94. Lampblack is a very finely divided form of carbon which is deposited on cold objects placed in the flames of burning oils. Oils are very rich in carbon, and to produce lampblack they are burned in a limited supply of air. When the dense smoke arising from them, which is mainly finely-divided carbon, is cooled, the carbon is deposited. Lampblack is one of the purest forms of amorphous carbon. It is used in making printer's ink and certain black paints.

Carbon is found also in all organic substances, both animal and vegetable. It is a constituent of all kinds of peat and humus, as well as of natural gas, petroleum, and asphalt. It exists in limestone, chalk, marble, and all other carbonates. In the air it is found as carbon dioxide. It forms more different compounds than any other element, and it is said that more than 100,000 carbon compounds have been prepared and analyzed.

95. Properties of Carbon. Notwithstanding their marked differences in appearance all these forms of carbon have some properties in common. They are insoluble in all ordinary liquids. They are tasteless and odorless. They cannot be melted. They can actually be changed one into the other, for both graphite and diamonds of microscopic size have been prepared artificially from amorphous carbon. At ordinary temperatures carbon is *inactive.* At high temperatures all forms can be made to burn, and the *product in each case is the carbon dioxide* (CO_2). Carbon is a strong reducing agent, and when heated with the oxide of a metal, for instance, will unite with the oxygen, leaving the metal free. When copper oxide is heated with charcoal the reaction may be written thus:

$$2\,CuO + C \rightarrow 2\,Cu + CO_2.$$

This method is commonly used in the extraction of metals from their ores. If the ore is an oxide, it is heated directly with carbon. If the metal is in the form of a sulphide it is first roasted to change it to the oxide, after which it is reduced by heating it with carbon. *The chief use of carbon, however, is as a fuel.*

96. Flames due to a Burning Gas. In Chapter V, combustion was defined as "the union of a substance with oxygen with the evolution of light and heat." In the everyday sense, however, combustion consists in the burning, or oxidation, of a material containing carbon; for all ordinary fuels, whether gaseous, liquid, or solid, contain carbon. A marked difference is noticeable in the manner in which the different fuels burn. Some of them burn with a flame and some do not. The gases all produce flames, and so do the liquid fuels. Of the solid fuels, wood and soft coal burn with a flame, while charcoal and coke do not. Anthracite gives a very feeble flame. Careful study of flames has shown that they *always consist of burning gases.* That this is true of a gaseous fuel is evident, but it is equally true of kerosene or other burning oils. In the case of the kerosene lamp, for instance, the gas which burns is produced from the oil which is drawn up the wick by capillary attraction and then volatilized by the heat of the flame.

In the case of the candle the heat first melts the wax, and the liquid thus formed is drawn up the wick and then converted into a gas that burns. That gas is actually formed during the burning of a candle can be shown by placing the lower end of a piece of glass tubing in the center of the flame; the gas passes up the tubing and can be ignited at the upper end *A* (Fig. 76). Flames are produced when wood and soft coal are burned, because both these materials contain volatile substances which are converted into gases by the heat. The method of preparing charcoal and coke is such that all these volatile substances are driven off. Charcoal and coke, therefore, burn without a flame because there is no volatile matter present to be converted into a gas. Anthracite contains very little volatile matter, and so the flaming during its burning is very slight.

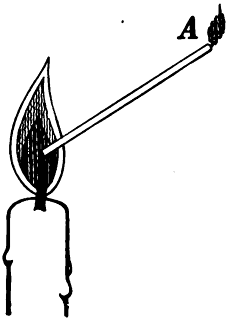


FIG. 76. — Ignition of gas from the inner zone of a candle flame.

97. Luminosity of Flames. There is a marked difference in the luminosity of flames, the variation depending partly on the gas itself and partly on the way in which it is burned. Hydrogen burns with a non-luminous flame; natural gas gives more light than hydrogen but not so much as coal gas; and acetylene burns with a more luminous flame than either natural gas or coal gas. It has been found that *when the combustion of a gas is complete the flame is always non-luminous.* The Bunsen burner used in the laboratory illustrates this point (Fig. 77). It is constructed with the idea of mixing the gas and the air in such proportions as to bring about the complete combustion of the gas. The gas enters at the base of the burner at *A* and is mixed with the air entering at the side holes *BB*, and the mixture of gas and air is burned at the top *C*. If the amount of air entering at *BB* is properly

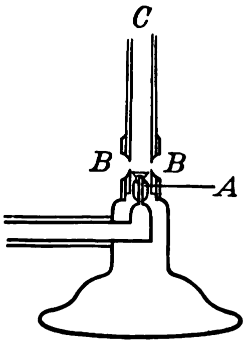


FIG. 77. — Section of a Bunsen burner.

adjusted, the flame will be *blue and non-luminous*. If the openings at the bottom of the burner are closed so that no air can enter, the flame becomes *yellow in color and luminous*.

The luminosity of the flame is easily explained. A piece of platinum wire placed in the blue flame of the Bunsen burner becomes white-hot and gives off light. If some fine iron dust is blown into it, the flame becomes momentarily luminous. The same effect may be produced with charcoal dust or fine table salt. In these experiments it is evident that the light comes from a solid substance which has been heated to a white heat or *to incandescence*. The same thing is true of all luminous flames. In the flame produced by any of the ordinary illuminants there is a place in the flame where the combustion is not complete. The heat decomposes some of the gaseous compounds and the carbon is set free. It is this very *finely divided carbon heated to incandescence* that gives off the light. A gas like acetylene, which is very rich in carbon, therefore, gives more light than one like natural gas, which is relatively low in carbon. In general, then, it may be said that light is produced by a solid substance which is heated to incandescence. In the Welsbach burner the light comes from the mantle, the material of which is heated by the blue flame of the Bunsen burner. In the limelight the lime is heated to incandescence by the oxyhydrogen flame, and in the electric bulb the filament is heated by the electric current; but the light in each case comes from the heated solid.

98. Structure of Flames. The luminous flame has several distinct parts, as can readily be seen in the flame of the candle. A vertical section of the candle flame is represented in Fig. 78. Around the wick there is a dark cone *A* filled with combustible gases formed from the melted wax, which do not burn because there is no oxygen present. It was from this cone that the gas was drawn in Fig. 76. Above the dark cone is the luminous part of the flame *B*. Here the oxygen is insufficient for complete combustion, but the temperature is sufficiently high to decompose some of the gas and liberate small particles of carbon. This liberated carbon heated by the burning gas makes the flame luminous. A piece of crayon held in this part of the flame will at once be coated with carbon. The exterior cone *C* is almost invisible, because here there is plenty of oxygen and combustion is complete, and all the carbon is burned to carbon dioxide. These three regions will be found in all illuminating flames, whatever their shape, as can be seen by carefully examining the flat flame of the ordinary gas burner or the flame of the kerosene lamp. In the non-luminous flame of the Bunsen burner two principal regions are easily distinguished, an inner cone *A* of unburned gas and an outer cone *B*, where the combustion is complete. The hottest part of the flame is just above the inner cone *A* (Fig. 79).

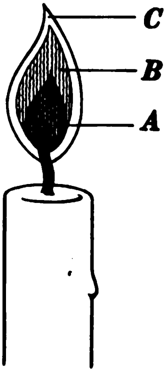


FIG. 78.—**A vertical section of candle flame showing the three zones.**

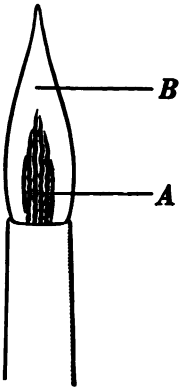


FIG. 79.—**A vertical section of the Bunsen flame showing the two zones.**

99. Kindling Temperature of Gases. Gases, like other substances, must be kept at their kindling temperature in order to burn. If they are cooled below the kindling temperature, the flame is extinguished. If a piece of wire gauze is pressed down on the flame of a Bunsen burner, the flame remains below the gauze, although the gas passes freely through it and escapes. If the gas is now extinguished and then relighted above the gauze, it will burn above but not beneath (Fig. 80). The explanation is that the gauze conducts away the heat rapidly enough to cool the gas below its kindling temperature.

100. Davy's Safety Lamp. The miner's safety lamp depends upon this principle. It is an oil lamp surrounded by fine wire gauze (Fig. 81). In a mine where there are explosive gases the lamp will continue to burn, and some of the gas may even enter the lamp and burn inside; but since the wire gauze prevents the gas on the outside from being heated to its kindling temperature, explosions are often prevented.

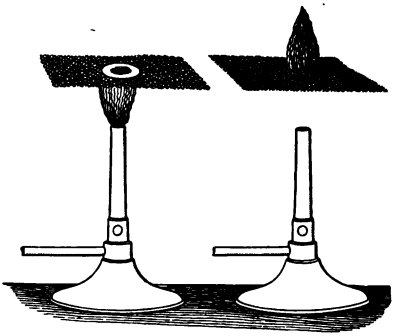


FIG. 80. — Showing how wire gauze cools the gas below the kindling point.

FIG. 81. — Davy's safety lamp.

EXERCISES

Ex. **66.** Heat some sawdust, a piece of cotton, a bit of bone, a piece of lean meat, and some sugar or starch in test tubes or in a covered iron dish. What residue do you get in these experiments? Hold a piece of crayon in the flame of a candle. What is the black coating on the crayon? What do these experiments show as to the distribution of carbon? Name some other substances containing carbon. Give the properties of carbon. What is formed when carbon burns? Tell what you can about the following forms of carbon: (1) diamond, (2) graphite, (3) wood charcoal, (4) animal charcoal, (5) coal, (6) lampblack.

Ex. 67. Mix a teaspoonful of copper oxide with an equal quantity of powdered charcoal and place it in a hard glass test tube. Arrange as in Fig. 82, allowing the end of the rubber tubing to dip into a bottle containing limewater. Heat the tube cautiously until gas ceases to be evolved, and remove tubing from the water. What has happened to the limewater? To the copper oxide? Explain the change in the copper oxide and write the reaction. What kind of agent is carbon? Is any commercial use made of this property of carbon?

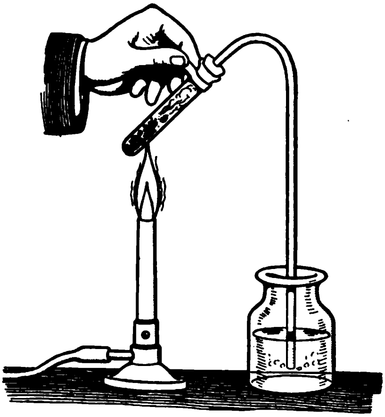


FIG. 82. — Heating charcoal and copper oxide and passing the resulting gas into limewater.

Ex. 68. Put half a test tube full of bone black into a small flask and pour in about two ounces of water to which have been added a few drops of indigo or litmus. Mix, heat gently for a few minutes, and filter. Has anything happened to the color? Name an industry in which animal charcoal is used as a decolorizer.

Fill a test tube half full of powdered wood charcoal. Add 2 cc. of a solution of hydrogen sulphide. Cork the tube securely and shake it thoroughly for some time. Let it stand for fifteen minutes; then remove the stopper and note whether the odor is less offensive. Is charcoal ever used as a deodorizer?

Ex. 69. Place some pieces of soft coal in a small porcelain crucible and connect it with the bowl of a clay pipe, making the connection tight with clay (Fig. 83). Heat in the Bunsen flame. Does anything escape through the stem of the pipe? Will the escaping material burn? When all the volatile matter has been expelled, examine the residue in the crucible. What is it? Explain how illuminating gas is manufactured. How is coke produced?

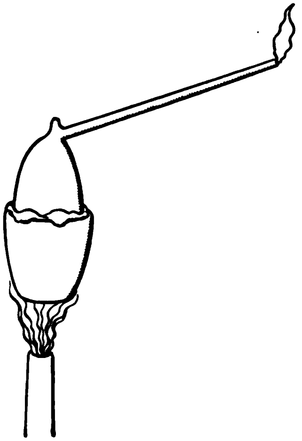


FIG. 83. — **Apparatus to illustrate the manufacture of illuminating gas and coke.**

Ex. **70.** Does the blue flame of the Bunsen burner give much light? Hold a piece of platinum wire in the flame. Is more light produced? Sprinkle a little charcoal dust or some fine salt in the flame. Does it make the flame more luminous? To what is the luminosity of the flame due? Why is the acetylene flame more luminous than that of natural gas? What makes the light in the Welsbach burner? Draw a diagram of the flame of a candle and indicate the different zones. What makes the flame of the candle luminous?

Ex. **71.** Do all substances burn with a flame? What substances produce flames? How can you show that the flame of a candle is due to a burning gas? Light a candle and allow it to burn a few minutes. Light a match, blow out the candle, and apply the match to the ascending smoke. Repeat, noting whether the candle can be lighted at a distance from the wick. Explain how this is possible. Why does wood burn with a flame while charcoal does not?

Ex. **72.** Press a piece of wire gauze halfway down on a Bunsen flame (Fig. 80). Does the flame pass through the gauze? Does any unburned gas pass through? Turn off the gas, then turn it on again and light it above the gauze (Fig. 80). Does the gas burn below the gauze? Explain the results in these two cases. How is the miner's safety lamp constructed? Explain how it prevents explosions. Do gases have a definite kindling temperature?

CHAPTER XII

CARBON COMPOUNDS

101. Carbon Burns to Carbon Dioxide. If a piece of charcoal is ignited and placed in a bottle containing oxygen, it will burn violently, throwing off a shower of sparks, and the bottle will be filled with a gas having a slightly pungent odor. Limewater placed in the bottle becomes milky, and a white precipitate settles out. The carbon has burned to carbon dioxide, which fact may be expressed thus:

$$C + 2\,O \rightarrow CO_2.$$

This behavior with limewater is a *test for carbon dioxide,* since no other gas acts in this way. If the products of combustion from a gas jet, from burning alcohol, or from the flame of a candle or kerosene lamp are collected by holding an empty wide-mouth bottle over the flame for a moment, and then tested with limewater, it will be found that the milkiness is produced in each case. In other words, *when any substance containing carbon is burned, the carbon is oxidized to carbon dioxide.*

102. Carbonic Acid. If a little water is added to a bottle in which charcoal has been burned, the carbon dioxide will dissolve in it. This water will *turn blue litmus paper red,* and has a very faintly sour taste, which suggests that an acid is present. When carbon dioxide is dissolved in water, it forms a weak, unstable acid according to the equation,

$$CO_2 + H_2O \rightarrow H_2CO_3.$$

This compound, H_2CO_3, is *carbonic acid.* It is so unstable that even at low temperatures it breaks up into carbon dioxide and water; thus:

$$H_2CO_3 \rightarrow CO_2 + H_2O.$$

(Compare with the reaction of sulphur dioxide and water (83).) The instability of carbonic acid makes it impossible to obtain it free from water. Many of its salts, which are called carbonates, are known. The carbonates are stable compounds and most of them are insoluble in water. Limestone, marble, and chalk are all calcium carbonate ($CaCO_3$), the salt formed by replacing the hydrogen of carbonic acid with metal calcium. Washing soda, or sodium carbonate (Na_2CO_3), is another well-known salt of carbonic acid.

103. Preparation of Carbon Dioxide. When a strong acid like sulphuric acid acts on a carbonate, the carbonic acid is set free; but since the latter is, for the most part, immediately decomposed into carbon dioxide and water, the reaction is usually written thus:

$$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O.$$

This reaction may be used for the preparation of carbon dioxide, but the more usual laboratory method is by the action of hydrochloric acid on marble, or ordinary limestone. Hydrochloric acid is used because its compound with calcium is soluble in water (Fig. 84), while calcium sulphate is not. The reaction is indicated thus:

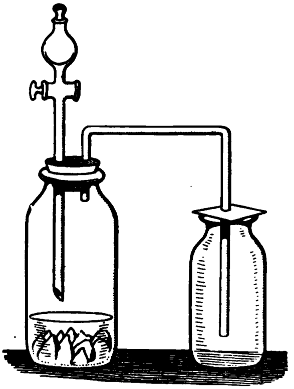


FIG. 84. — **Production of carbon dioxide from limestone and hydrochloric acid.**

$$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + CO_2 + H_2O.$$

104. Properties of Carbon Dioxide. Carbon dioxide is a colorless gas with a slightly pungent odor and acid taste. It is one and one half times heavier than air. It can be poured from one vessel to another. At ordinary temperature and pressure water dissolves its own volume of carbon dioxide. Under increased pressure water dissolves much more of the gas, and as the pressure is released the gas escapes. Soda water is made by forcing carbon dioxide into water under pressure, and the escape of the gas as the pressure is released accounts for the *effervescence* and frothing when it is drawn from the fountain or a siphon bottle (Fig. 85). Many mineral waters, as well as manufactured beverages, sparkle and effervesce for the same reason. Carbon dioxide is somewhat easily liquefied and is sold in large quantities in steel cylinders. Liquid carbon dioxide is used in the manufacture of soda water and to produce very low temperatures. Many of the small fire extinguishers (Fig. 86) contain baking soda and sulphuric acid so arranged that they can be mixed at the moment needed. The mixture of water and CO_2 under pressure will often put out a small blaze and prevent a serious fire. Air containing 3 to 4 per cent of carbon dioxide will extinguish small flames.

FIG. 85. — Siphon bottle used to hold carbonated water.

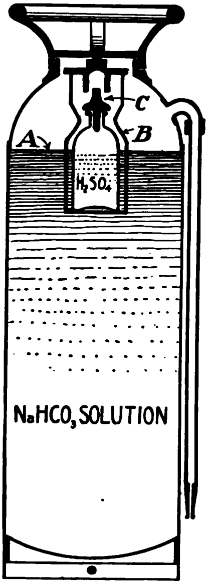


FIG. 86. — Section of a fire extinguisher.

Carbon dioxide will not burn because it is itself *the product of the complete combustion of carbon.* Carbon dioxide will not support combustion nor sustain life. It is not poisonous, but animals placed in it die of suffocation. Water containing carbon dioxide will dissolve many substances which are but slightly soluble in pure water. This property of carbonated water is due to the presence of *carbonic acid,* which is formed when carbon dioxide dissolves in water. All soil water contains carbonic acid, which is an important factor in changing the insoluble constituents of the soil into soluble forms.

105. Carbon Dioxide and Plant Life. It has been shown that carbon dioxide comprises about .03 per cent to .04 per cent of the atmosphere, or three to four parts in ten thousand. This carbon dioxide is the sole source of the carbon found in green plants, and, since all animals live directly or indirectly upon plants, it is the source of all the carbon found in both animal and vegetable tissues. Green plants have the power of abstracting carbon from carbon dioxide and uniting it with other substances to form the various complex compounds found in the plants. This power of the green plant is dependent upon the green coloring matter, or chlorophyll, which is found in the leaves. The first visible effect of the action of chlorophyll is the presence of starch in the leaves. The chlorophyll apparently brings about a reaction between carbon dioxide and water, which may be represented by the following equation:

$$6\,CO_2 + 5\,H_2O \rightarrow \underset{\text{starch}}{C_6H_{10}O_5} + 12\,O.$$

That green leaves decompose carbon dioxide and liberate the oxygen can be shown by a simple experiment. A quantity of water is charged with carbon dioxide by running a current of the gas from a generator through it for a few minutes. Some sprigs of mint, water cress, or some other plant are placed in a glass cylinder and covered with carbonated water. A funnel and a test tube filled with water are arranged, as shown in Fig. 87, so as to collect any gas which may be formed. The apparatus is placed in the strong sunlight, and after a short time bubbles of gas will arise and replace the water in the test tube. When sufficient gas has collected, the test with a glowing splint will show that it is oxygen.

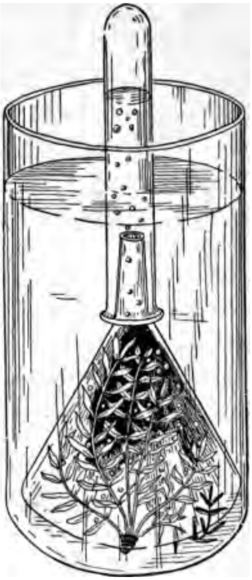


FIG. 87. — **The decomposition of carbon dioxide and liberation of oxygen by plants.**

Sunlight is necessary to furnish the plant with the energy required to decompose the carbon dioxide. When carbon burns to carbon dioxide, a large amount of heat is given off. This heat is known as the *heat of formation*. Before such a compound as carbon dioxide can be decomposed, an amount of energy must be provided which is equivalent to the heat energy which was given off when the compound was formed; or, in other words, an *amount of energy equivalent to the heat of formation of the compound.* The plant derives this energy from the sunlight, and consequently the formation of starch in the plant does not go on at night. It will thus be seen that without sunlight all life would cease, for every living thing is dependent either directly or indirectly upon the decomposition of carbon dioxide by plants (340). The process by which plants utilize the carbon of carbon dioxide to form starch is called *photosynthesis* (341).

106. Amount of Carbon Dioxide in the Atmosphere. The percentage of carbon dioxide in the atmosphere is so small that it might be feared that the supply would soon be exhausted. So great is the bulk of the atmosphere, however, that it has been calculated that the air contains not less than 3,400,000,000,000 tons of carbon dioxide. This amounts to 28 tons over each acre of the earth's surface.

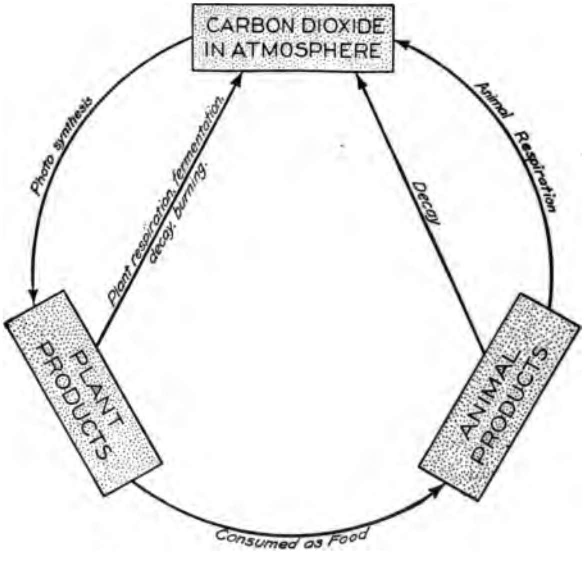


FIG. 88. — The cycle of carbon in nature.

107. Formation of Carbon Dioxide in Nature. The supply of carbon dioxide in the atmosphere is being constantly renewed in several ways. The burning of fuels of all kinds results in the production of carbon dioxide, as does also the decay of all organic matter. Fermentations, such as take place in wines and cider, and in breweries and distilleries, give rise to carbon dioxide. It is also given off in the vicinity of volcanoes and from mineral springs. It is exhaled by the breathing of all animals as well, for it is the product of the slow combustion of carbon in the animal body. Carbon dioxide sometimes accumulates at the bottom of wells, in silos, and in mines, since it is often formed in such places more rapidly than it can be removed by diffusion. Many deaths have occurred from suffocation in such places. Carbon dioxide is called *choke damp* by the miners. The production of carbon dioxide in the various ways mentioned is so nicely balanced by the decomposition of this gas by green plants, that the amount of carbon dioxide and of oxygen in the air scarcely varies. The cycle of carbon in nature is indicated in Fig. 88.

108. Carbon Monoxide. When a substance containing carbon is burned *in an insufficient supply of air, carbon monoxide* (CO) *is formed.* If carbon dioxide is passed over highly heated carbon, a reaction takes place which is represented by the following equation:

$$CO_2 + C \rightarrow 2\,CO.$$

Carbon monoxide is formed during the burning of hard coal in a stove or grate. At the lower part of the fire, where there is free access of air, the carbon burns to carbon dioxide, but as it passes up through the heated coal the dioxide is partially reduced to carbon monoxide. When the monoxide escapes from the top, it again combines with oxygen and burns with the blue flame always noticed above a mass of burning hard coal. Pure carbon monoxide is a colorless, tasteless, and odorless gas, which burns with a pale blue flame. *It is exceedingly poisonous.* It is the most dangerous gas given off from coal stoves, and great precaution should be taken to prevent its escape into the room.

109. Water Gas. In the manufacture of coal gas it is customary to take advantage of the fact that highly heated carbon will decompose water. After all the volatile matter has been driven off from coal, and while the coke is still very hot, steam is turned into the retorts with the result that the following reaction takes place:

$$H_2O + C \rightarrow CO + H_2.$$

This mixture of hydrogen and carbon monoxide is known as *water gas*. Only a limited quantity of the gas can be made, as the coke is soon cooled below the point at which it will decompose water. As both hydrogen and carbon monoxide burn with a non-luminous flame, water gas cannot be used as an illuminating gas unless it is first enriched by the addition of some petroleum product high in carbon. Owing to its high percentage of carbon monoxide, water gas is *very poisonous.*

110. Carbon Bisulphide. Carbon forms one important compound with sulphur. The two elements are made to combine in an electric furnace, the resulting compound being carbon bisulphide, CS_2. Commercial carbon bisulphide is a yellow liquid with an offensive odor. It is poisonous, volatile, and very inflammable. The equation for its combustion is

$$CS_2 + 6\,O \rightarrow CO_2 + 2\,SO_2.$$

Carbon bisulphide is insoluble in water. It dissolves rubber, gums, fats, camphor, and sulphur (60). The common rubber cement is a solution of rubber in carbon bisulphide. Carbon bisulphide is used as an insecticide and to exterminate burrowing animals such as moles, woodchucks, and gophers. As its vapor is much heavier than air it readily sinks to the bottom of the burrow.

111. Other Compounds of Carbon. It has been stated that more than 100,000 compounds of carbon have been prepared. Most of these compounds are of animal or vegetable origin, and the study of these numerous and complex compounds is commonly called *organic chemistry*. It was formerly thought that these compounds could be produced only by life processes, but many of them have been produced artificially. It is more convenient, however, to deal with this subject after the chemistry of the other common elements has been studied. The more common and important of the carbon compounds will be discussed in Part II of this text.

EXERCISES

Ex. 73. Ignite a piece of charcoal and place it in a bottle containing oxygen. Note the odor of the gas formed. Place a little limewater in the bottle. What happens to the limewater? What is the gas formed by burning carbon? Hold an empty bottle over the flame of a candle and test it with limewater. What is the result? Take some limewater home with you and test any flames for carbon dioxide.

Ex. 74. Place a little water in the bottom of a wide-mouth bottle and burn a piece of charcoal in the bottle. Shake the bottle and test the water with litmus paper. Is an acid formed? Write the equation. What happens when the solution is warmed? Compare with the reactions for sulphur dioxide. Can carbonic acid be prepared in the pure state? Are its salts (the carbonates) stable? Name a common carbonate.

Ex. **75.** Arrange apparatus as in Fig. 84 and place several pieces of marble or limestone in the bottle and cover with water. Pour dilute hydrochloric acid in the thistle tube and collect the carbon dioxide by downward displacement as shown in the figure. What is the appearance of the gas? What is its odor and its taste? Light a candle, place it in a tumbler, and pour a bottleful of carbon dioxide over it (Fig. 89). What happens to the candle? Is CO_2 heavier than air? Place a bottle of air mouth downward over a bottle of carbon dioxide. After ten minutes test the contents of the upper bottle for CO_2. What is the result? What effect does pressure have on the solubility of the gas in water? What use is made of this fact? What use is made of liquid carbon dioxide? How do the small fire extinguishers work? Why will carbon dioxide not burn? Is it poisonous? Why do animals die when placed in it? Why does water containing carbon dioxide dissolve some minerals which are insoluble in pure water? How many substances can you find at home which are carbonates or contain carbonates? Test them with hydrochloric acid and note whether CO_2 is given off.

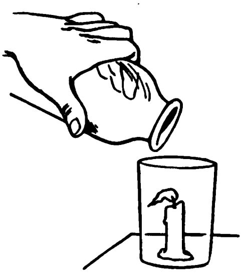


FIG. 89. — **Extinguishing a candle by pouring carbon dioxide gas over it.**

Ex. **76.** Fill a tall glass cylinder with water and cause carbon dioxide to bubble through it for a few minutes. Place several sprigs of mint or water cress in the cylinder and arrange a funnel and test tube as shown in Fig. 87. Place the apparatus in strong sunlight and when sufficient gas collects in the test tube test for oxygen with a glowing splint. How do green plants obtain their carbon? Is chlorophyll necessary for this process? Why is sunlight necessary to enable the plant to use the carbon of carbon dioxide? What is heat of formation? How much carbon dioxide does the atmosphere contain? Is there any danger of the supply being exhausted by plants? How is the supply of CO_2 in the air renewed? Why is it dangerous to go into a well or a silo that has been closed tightly for some time? What name do the miners give to carbon dioxide?

Ex. 77. What other oxide of carbon is known? Explain the formation of this oxide in the coal stove. Reaction? Why should coal stoves and furnaces be gas tight? How is water gas made? Give reaction. Is it a good illuminating gas? How can it be improved as an illuminant? What important compound does carbon form with sulphur? Reaction? In what experiment did you use this substance? For what purposes is carbon bisulphide used? If there is an ant hill at home try the following experiment. With a stick make a hole an inch or two in diameter and a foot deep in the ant hill. Pour in two ounces of carbon bisulphide and cover the hole with earth. Place a piece of carpet or a blanket over the ant hill. Examine after twelve hours and report.

CHAPTER XIII

LIMESTONE AND OTHER CALCIUM COMPOUNDS

112. Limestone. One of the most widely distributed substances in nature is *limestone.* It is most familiar as the ordinary limestone used for building purposes; but marble, chalk, coral, marl, and shells are identical in composition with limestone. They all consist largely of the compound made by the combination of the metal calcium with carbonic acid; namely, calcium carbonate, $CaCO_3$. In marble the calcium carbonate exists as a mass of minute crystals. Chalk and most limestone are not crystalline, but often show by their structure that they have been derived from shells. The transparent crystals of *calcite* and *Iceland spar* (Fig. 91) are very pure forms of cal-

FIG. 90. — A limestone cliff.

cium carbonate. Calcium carbonate is insoluble in water. It is decomposed by most acids, with the result that carbon dioxide is given off and the calcium salt of the acid is formed. (See 103.)

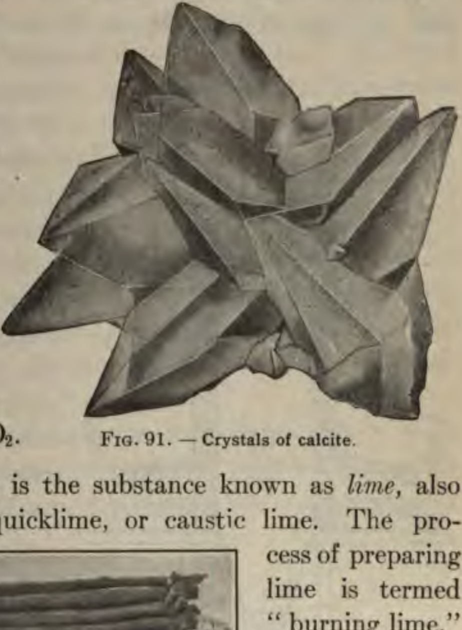


FIG. 91. — Crystals of calcite.

113. Manufacture of Lime. When calcium carbonate is strongly heated, carbon dioxide is driven off and calcium oxide remains, thus:

$CaCO_3 \rightarrow CaO + CO_2$.

Calcium oxide, CaO, is the substance known as *lime,* also called burnt lime, quicklime, or caustic lime. The process of preparing lime is termed "burning lime," which is a misnomer, as burning consists in the uniting of a substance with oxygen, while this process is a decomposition and not an oxidation. The "burning" of

FIG. 92. — A homemade limekiln.

lime is one of the oldest of chemical processes and has been carried on for at least fifty centuries.

In the older method of preparing lime a fire of wood or coal is made at the bottom of the limekiln, which is a shaft or chimney quite commonly built in a hillside. When the limestone which is placed on the fuel is completely burned, it is removed and the kiln is refilled. At the present time kilns (Fig. 93) are so built that the operation is continuous, limestone being added at the top of the kiln and the lime being removed from the bottom as fast as it is formed. Lime is most familiar in the form of lump, or builder's, lime. It is often ground to a powder and sold as ground lime.

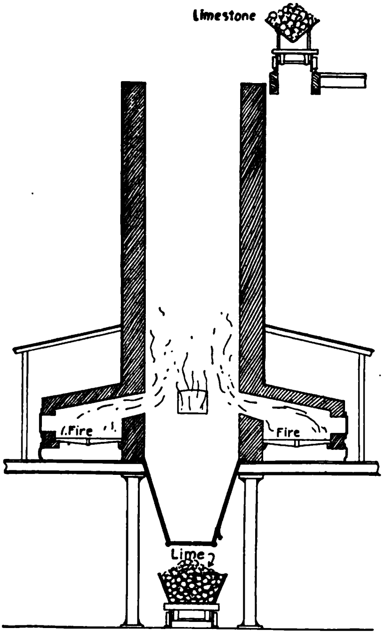


Fig. 93. — A modern limekiln.

114. Slaked Lime. When lime is sprinkled with water, it becomes very hot, swells, and finally crumbles to a white powder. This process is called slaking the lime, and the white powder is known as *slaked lime* or *hydrated lime.* The chemical name for it is *calcium hydroxide.* The change may be represented thus:

$$CaO + H_2O \rightarrow Ca(OH)_2.$$

Calcium hydroxide, $Ca(OH)_2$, is somewhat soluble in water, and this solution is called limewater. It will thus be seen that calcium oxide behaves toward water in much the same way as do the oxides of sulphur and carbon; that is, it forms a chemical compound with the water. But when sulphur trioxide, for instance, is added to water, the solution has a sour taste and it turns blue litmus paper red. Limewater, on the other hand, has an astringent, or alkaline, taste, and does not change the color of blue litmus paper. On the contrary, if the paper which was turned red by the acid is placed in limewater, *the blue color will be restored.* It will be found that there are other oxides which behave in the same way. Evidently, then, not all oxides form acids with water, but some form compounds with properties the opposite of acids.

Limewater is used in medicine, and as previously mentioned is used in the chemical laboratory as a test for carbon dioxide, with which it forms the insoluble calcium carbonate, the equation being

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O.$$

When considerable calcium hydroxide is suspended in water, the mixture is called *milk of lime.* Ordinary whitewash is thin milk of lime.

Mortar is a thick paste made by mixing slaked lime and sand. It sets or hardens partly owing to the loss of water by evaporation, but also because carbon dioxide is absorbed and the calcium hydroxide is changed into the carbonate. The same change takes place in whitewash when it is spread on a wall. The hardening of plaster may be hastened by burning charcoal in the room so as to increase the amount of carbon dioxide in the air.

115. Air-slaked Lime. When lime is exposed to the air, it absorbs moisture and slakes. It also absorbs carbon dioxide from the air and changes to the carbonate:

$$CaO + CO_2 \rightarrow CaCO_3.$$

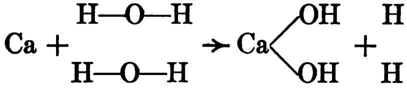
Lime that has undergone this change to the hydroxide and carbonate is said to be air-slaked and is of no value for making mortar. Fires are known to have been caused by the heat generated by the action of moisture on lump lime. The agricultural uses of lime are discussed in Chapter LV.

116. Cement. Portland cement is made by mixing limestone and clay, in the proper proportions, and heating them strongly until the mass begins to melt. The *clinker* formed in this way is ground to fine powder, which is the cement. Occasionally an impure limestone is found which contains the right quantity of clay and sand, and this is burned directly to make the so-called *natural cement.* When cement is moistened it sets to a hard, stone-like mass. Mixed with sand and stone it forms concrete, which is rapidly taking the place of stone in many building operations. The chemical changes which take place in cement are not understood.

117. Calcium. The element calcium is the fifth most abundant element in nature, but *it is never found uncombined.* It is seldom seen outside of chemical laboratories. It is a silver-white metal, soft enough to be cut with a knife. It will burn to the oxide ($Ca + O \rightarrow CaO$), which is the same compound as that formed by heating the carbonate. *It decomposes water at ordinary temperatures,* forming the hydroxide and giving off hydrogen, thus:

$$Ca + 2\ H_2O \rightarrow Ca(OH)_2 + 2\ H.$$

When steam was passed over hot zinc, as was seen in Chapter II, the metal extracted the oxygen from the water to form the oxide and set free both atoms of hydrogen. When calcium acts upon the water at ordinary temperatures, it liberates *only one atom* of hydrogen from the molecule of water, and the oxygen and the remaining hydrogen atom unite with the metal. The change may be represented thus:



The calcium atom, being bivalent, has *replaced one hydrogen atom from each of two molecules of water.* The group OH, which may be said to be water with one hydrogen atom removed, is of common occurrence in chemical compounds and has been named *hydroxyl.*

118. Calcium sulphate occurs in large quantities and is widely distributed in the substance known as gypsum, which has the formula, $CaSO_4 \cdot 2\ H_2O$. The two molecules of water set off by the period in the formula represent the *water of crystallization.* Although the water of crystallization can be driven off by heat, there is good reason to believe that it is in *chemical combination* with the rest of the compound. Water of crystallization is not apparent as moisture. It is in some way essential to the crystalline form, in which it is always present in a definite proportion. The molecule of gypsum always crystallizes with two molecules of water.

Gypsum occurs as white masses or as transparent crystals. It is slightly soluble in water, one part dissolving in 400 parts of water. When heated it loses part of its water of crystallization and is known as *plaster of Paris.* When water is added to plaster of Paris, the lost water is taken up again, and the material is changed back to gypsum and sets as a hard, compact mass. Plaster of Paris is used in coating walls, in making stucco, and in making casts and reproductions of statuary and small objects. If the gypsum is heated too strongly, the plaster is spoiled and will not set.

119. Hardness of Water. In Chapter II mention was made of *temporary* and *permanent hardness* of water, and it is now possible to explain these phenomena. *Temporary hardness* is caused for the most part by the *presence of calcium carbonate* in the water. Calcium carbonate will not dissolve in pure water, but will do so to a limited extent in water containing carbon dioxide. This is due to the formation of *acid calcium carbonate* $Ca(HCO_3)_2$, thus:

$$CaCO_3 + H_2CO_3 \rightarrow Ca(HCO_3)_2.$$

In acid calcium carbonate it is supposed that the atom of calcium *replaces one hydrogen atom from each of two molecules of carbonic acid.*

FIG. 94. — A cone built up from deposits made by carbonated spring water.

As most natural waters contain carbon dioxide, this compound is generally present in water. It is very unstable and when heated is decomposed into ordinary calcium carbonate, carbon dioxide, and water:

$$Ca(HCO_3)_2 \rightarrow CaCO_3 + CO_2 + H_2O.$$

This reaction explains why *boiling the well water causes a deposit of calcium carbonate and makes the water less hard.* On a large scale it is not practicable to boil the water to soften it, and the same result is brought about by adding just sufficient slaked lime to convert the acid carbonate into the normal carbonate, as shown in the following equation:

$$Ca(HCO_3)_2 + Ca(OH)_2 \rightarrow 2\,CaCO_3 + 2\,H_2O.$$

Hardness of water is not entirely due to calcium compounds. Magnesium (Chapter XXII) is usually associated with calcium in nature, and part of the hardness of water is usually due to magnesium compounds, which act in much the same way as do the compounds of calcium.

Permanent hardness of water is due to the presence of gypsum in solution, a compound which is not precipitated by boiling. To remove calcium sulphate ordinary washing soda (sodium carbonate) is added, and the following reaction takes place:

$$CaSO_4 + Na_2CO_3 \rightarrow CaCO_3 + Na_2SO_4.$$

The calcium carbonate being insoluble is precipitated and the sodium sulphate (Na_2SO_4) which stays in solution is not objectionable. Such a chemical change as the one represented in this reaction is called a *double decomposition.* It will be noticed that metals have changed places in the salts. The change might be represented thus —

$$\begin{tikzcd} Ca \arrow[dr] & SO_4 \\ Na_2 \arrow[ur] & CO_3 \end{tikzcd}$$

It is a general rule of chemistry that when two compounds in solution are mixed, their component parts rearrange themselves to form the most insoluble compound possible, and of course the insoluble compound is then precipitated. Calcium carbonate, being the most insoluble compound possible in this mixture, is precipitated, and the sodium sulphate, being very soluble, remains in solution.

In the softening of water for technical purposes the water is carefully analyzed and the proper amount of calcium hydroxide is first added to precipitate the carbonates. After a short time the sodium carbonate is added to remove the other calcium and magnesium salts. The household methods for softening water are discussed in Chapter XLVIII.

EXERCISES

Ex. **78**. Place two or three thin chips of marble on a wire gauze and heat them ten minutes in the full flame of the burner. Compare with an unheated bit of marble. What is the composition of the marble? What change took place upon heating? Write the reaction. What substances can you name that are calcium carbonate? Can you name anything in your home that is composed of calcium carbonate? How is lime made commercially?

Ex. **79**. Lay one of the heated chips of marble from the above experiment on a piece of *red* litmus paper and on another piece of the same paper place a bit of unchanged marble. Moisten both pieces and describe what happens. Is there any difference in the effect of the two pieces of stone on the litmus paper?

Ex. **80**. Place a lump of quicklime in a beaker and moisten with hot water. Watch and describe the result. Write the reaction for the change in the lime. Fill the beaker with water and stir. When the solid settles, decant the clear liquid into a clean bottle. Is the calcium hydroxide soluble? Test a portion of the limewater with blue and red litmus papers. What is the effect? How does the compound of calcium oxide and water compare with the compound of sulphur dioxide and water? What is meant by milk of lime? By whitewash? What is mortar? What makes mortar set? What is meant by air-slaked lime? Is it of any value for building purposes? Tell what you can about the manufacture and uses of Portland cement. What are natural cements? Report on anything at home that is built with Portland cement or concrete.

Ex. **81**. Describe the element calcium. What compound is formed when it burns? Give the equation for the reaction of calcium on water. (Note to the teacher. Metallic calcium may be obtained at a moderate cost from some of the supply houses. It will be well to perform the experiment of liberating hydrogen from water by means of metallic calcium if possible.) Does calcium replace all the hydrogen of water? What name is given to the group OH in chemistry? How many hydroxyl groups unite with one atom of calcium?

Ex. **82.** Examine crystals of calcium sulphate. What is the formula? The common name? What happens to gypsum when heated? Mix plaster of Paris to a thick paste with water. Rub a little vaseline or lard over the surface of a button and press the button into the surface of the moist plaster. When the plaster has hardened, remove the button and describe the result. What made the plaster of Paris set? What uses can you give for plaster of Paris?

Ex. **83.** Dilute a little limewater with an equal volume of distilled water. Pass carbon dioxide into the solution. Does a precipitate form? . Continue the experiment and note if the precipitate disappears. Write two equations, one for each change. Boil a portion of this liquid. What happens? Explain the reappearance of the precipitate. Boil some well water from home. Does a precipitate form? Show how this experiment explains temporary hardness of water. To another portion of the solution formed above by passing carbon dioxide into limewater add a fresh portion of limewater. What result do you obtain? How can lime be used to remove temporary hardness of water?

Ex. **84.** To the well water which has been boiled to remove the temporary hardness add sodium carbonate (washing soda). Is a precipitate formed? Explain how permanent hardness is removed from water. Give the reaction. What is meant in chemistry by a double decomposition?

CHAPTER XIV

SALT: CHLORINE AND SODIUM

120. SALT is familiar to everyone as a white solid which, upon close examination, is seen to consist of cubical crystals. It is widely distributed. It is found in solution in sea water to the extent of 3.5 per cent, and in some countries all the table salt is obtained from this source. The water of Great Salt Lake contains over 20 per cent of salt. The substance occurs in many places as rock salt, and is often mined in large blocks and sold in this form or purified for table use. Much of the salt of commerce is obtained by drilling wells into the salt deposits, and introducing water, which dissolves the salt, forming a brine which is pumped out and evaporated.

FIG. 95. — **The making of salt by the evaporation of sea water in France.**

The appearance and taste of salt are familiar. Water dissolves 35 per cent of its own weight of salt, the substance being nearly as soluble in cold water as in hot, in which respect it differs from most solids. When it is pure it remains dry on exposure to the air, but it usually contains a little calcium or magnesium chloride and hence attracts moisture. In addition to its well-known uses, it is the chief source of the compounds of sodium and chlorine.

121. Action of Sulphuric Acid on Salt. When sulphuric acid is poured upon common salt, a gas is given off which has a strong, irritating odor. It fumes strongly in the air, and a burning candle thrust into it is immediately extinguished. This gas is extremely soluble in water, one volume of water dissolving over 500 volumes of the gas. Formerly this gas was called *spirit of salt* because it was prepared from salt.

The solution of the gas in water is strongly acid. It is *very sour and turns blue litmus paper red.* Its acid properties were early recognized, and it received the name of muriatic acid. This solution will be found, upon comparison, to be identical with the *hydrochloric acid* of the laboratory, which is in reality a solution of hydrochloric acid containing about 40 per cent of the gas.

122. Composition of Hydrochloric Acid. If the solution of hydrochloric acid is poured upon zinc, *hydrogen will be evolved* just as it is when sulphuric acid and zinc are used. It is fair to assume, therefore, that hydrochloric acid contains hydrogen as one of its elements.

If another portion of strong hydrochloric acid solution is placed in an apparatus like that shown in Fig. 96, with a quantity of manganese dioxide, and the flask is gently heated, a gas is evolved which has a yellow-green color and an irritating odor. This gas was discovered by Scheele in 1774. About 1810 the English chemist Sir Humphry

Davy proved it to be an element and named it *chlorine* because of its color. This gas also evidently came from the hydrochloric acid, since the manganese dioxide and water alone will not produce it.

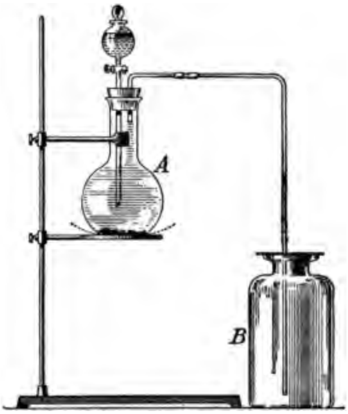


FIG. 96. — **Apparatus for the production of chlorine.**

If a mixture of chlorine and hydrogen is exposed to strong sunlight or the light from burning magnesium, it explodes, and the *hydrogen and chlorine combine to form hydrochloric acid gas.* This method of *synthesizing* hydrochloric acid shows that it contains the two elements hydrogen and chlorine and nothing else. If the combination were brought about in a eudiometer (Fig. 54), it would be found that exactly *equal volumes of the two gases combine.* As chlorine is 35.46 times as heavy as hydrogen, it follows that hydrochloric acid is composed of 1 part by weight of hydrogen to 35.46 parts by weight of chlorine. Its molecular weight, therefore, is 1 + 35.46 or 36.46. The formula assigned to it is HCl. It is an example of a very important acid that contains no oxygen.

123. Chlorine is irritating to the lining of the nose and throat and if breathed in large quantities causes inflammation. It is 2.5 times heavier than air, and hence is collected by downward displacement. Chlorine is one of the most active elements and is *never found in the free state.* It dissolves in water, the solution being known as chlorine water.

This solution is frequently used as a substitute for the gas. If the solution is placed in the sunlight, oxygen is liberated and hydrochloric acid is formed:

$$H_2O + Cl_2 \rightarrow 2\,HCl + O.$$

Chlorine is a powerful bleaching agent. This property probably depends upon the above reaction, the bleaching being due to the liberated oxygen, for the chlorine does not act upon colored fabrics unless they are moist. Chlorine combines directly with many metals, forming chlorides:

$$Cu + 2\,Cl \rightarrow CuCl_2.$$

The chlorides may also be formed by the action of hydrochloric acid on the metal or on an oxide of the metal:

$$Zn + 2\,HCl \rightarrow ZnCl_2 + 2\,H;$$
$$CaO + 2\,HCl \rightarrow CaCl_2 + H_2O.$$

124. Bleaching Powder. Large quantities of chlorine are needed in the bleaching industries, but as it is inconvenient to handle or transport the free gas it is stored in the form of *bleaching powder*, sometimes called *bleach* or improperly *chloride of lime.* It is made by passing chlorine over slaked lime:

$$Ca(OH)_2 + 2\,Cl \rightarrow CaOCl_2 + H_2O.$$

The formula $CaOCl_2$ is the one usually given for bleaching powder. When needed, the chlorine can be again obtained from the bleaching powder by treating it with sulphuric acid:

$$CaOCl_2 + H_2SO_4 \rightarrow 2\,Cl + CaSO_4 + H_2O.$$

In addition to its use in providing chlorine for bleaching, this powder is valuable as a disinfectant. It is slowly decomposed by the carbon dioxide of the atmosphere, and the chlorine is liberated:

$$CaOCl_2 + CO_2 \rightarrow CaCO_3 + 2\,Cl.$$

The disinfecting properties are probably due to the liberation of oxygen produced by the action of the chlorine on water.

125. Sodium. It is comparatively easy, as has been shown, to determine that chlorine is one of the elements found in salt, but there is no method suited to the small laboratory that will show what else is present in salt. If, however, the salt is fused and an electric current is passed through it, chlorine will be given off at the positive pole, while at the negative pole a new substance will be found. It is a soft solid which has a silver-white metallic luster, and becomes covered with a coating of white material when exposed to the air. It is the metal known as *sodium* (Na). If a bit of sodium is warmed and placed in a bottle full of chlorine, it burns with a dazzling yellow light and a white powder is formed which can be identified as salt. Shavings of cold sodium thrown into a jar of chlorine are slowly converted into a white mass of salt. Common salt, then, is the compound formed by the union of the metal sodium and the gas chlorine. It is sodium chloride (NaCl) and is evidently the *sodium salt of hydrochloric acid,* for it yields that acid when acted upon by sulphuric acid, the equation being:

$$2\,NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2\,HCl$$

Sodium is *very active* chemically and combines readily with many other elements, especially with chlorine and oxygen. It is kept in coal oil, as it absorbs oxygen and moisture from the air and quickly tarnishes. It decomposes water more vigorously than calcium does, liberating one half the hydrogen and forming sodium hydroxide:

$$Na + H_2O \rightarrow NaOH + H.$$

The water containing *sodium hydroxide* ($NaOH$) has a soapy feel when rubbed between the fingers and it exhibits the property of *turning red litmus paper blue* in a more marked way than does limewater. When burned in the air sodium is changed to sodium oxide ($2\,Na + O \rightarrow Na_2O$), which added to water gives the hydroxide:

$$Na_2O + H_2O \rightarrow 2\,NaOH.$$

126. Sodium hydroxide is better known under the name of *caustic soda* or *soda lye.* It is a white solid which readily absorbs water and carbon dioxide if exposed to the air. It has a very corrosive effect on animal and vegetable tissues. The chemically pure hydroxide is usually cast into sticks. Sodium hydroxide is used in many industries, especially in the manufacture of hard soaps. The material sold in cans as *potash* or *lye* is usually crude sodium hydroxide. It is sometimes produced by passing the electric current through strong brine, a process which causes hydrogen and chlorine to be given off and which allows the hydroxide to remain in solution. It is also made by boiling sodium carbonate with milk of lime:

$$Na_2CO_3 + Ca(OH)_2 \rightarrow 2\,NaOH + CaCO_3.$$

The solution of sodium hydroxide is separated from the insoluble calcium carbonate and concentrated by heating in iron kettles.

127. Sodium sulphate ($Na_2SO_4 \cdot 10\,H_2O$) is commonly known as Glauber's salt. It is produced when sulphuric acid acts upon common salt (121). It forms large crystals which are efflorescent; so the commercial article usually contains some white powder. It is used in medicine and in the manufacture of washing soda and of glass.

128. Sodium sulphite ($Na_2SO_3 \cdot 7\ H_2O$) was mentioned in Chapter X. It is prepared by the action of sulphur dioxide on sodium hydroxide. The equation is

$$SO_2 + 2\ NaOH \rightarrow Na_2SO_3 + H_2O.$$

When treated with an acid, sodium sulphite yields sulphur dioxide (86).

129. Sodium carbonate ($Na_2CO_3 \cdot 10\ H_2O$) is the substance known as *washing soda* or *sal soda.* It crystallizes in large crystals which are strongly efflorescent. When the sulphite and the carbonate are dried until most of the water of crystallization is driven off they are said to be *anhydrous.* The solution of sodium carbonate is slightly alkaline and will turn red litmus paper blue. It is used to soften water and to make soap, and in the manufacture of glass and many chemical reagents.

130. Sodium bicarbonate ($NaHCO_3$) or *baking soda* is carbonic acid with only one of the hydrogen atoms replaced by sodium. It can be made by passing carbon dioxide through a strong solution of sodium carbonate; or in other words, by the action of carbonic acid on sodium carbonate. The equation is:

$$Na_2CO_3 + H_2CO_3 \rightarrow 2\ NaHCO_3.$$

As the bicarbonate is much less soluble than the carbonate, it settles out. When heated it gives off carbon dioxide and water, and changes back to the normal carbonate:

$$2\ NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O.$$

When mixed with any acid material, such as sour milk or tartaric acid, it gives off carbon dioxide. This property accounts for its use in cooking, the liberated carbon dioxide being the substance that makes bread, pastry, or cake light (427).

131. Test for Sodium. To test for sodium dip a platinum wire into the substance to be tested and place it in the non-luminous flame (Fig. 97). A deep yellow color given to the flame shows the presence of sodium. As sodium chloride is present nearly everywhere, precaution must be taken to prevent the accidental introduction of this material when testing for sodium.

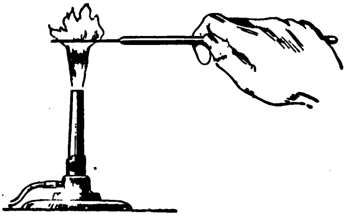


FIG. 97. — **Flame test for sodium.**

132. Test for Hydrochloric Acid or a Chloride. Add to the solution of the substance a few drops of the laboratory solution of silver nitrate. A white curdy precipitate, or a milkiness, *which does not dissolve in nitric acid,* is proof of the presence of hydrochloric acid or a chloride.

EXERCISES

Ex. 85. Tell what you can about the distribution of salt in nature. How is salt prepared for use? To what extent is salt soluble? Is it any more soluble in hot water than in cold? Is that true of most soluble solids? Why does ordinary table salt become moist in the air? Will pure salt remain dry?

Ex. 86. Place some salt in the flask (*A*) of an apparatus like Fig. 96 and add sulphuric acid. Gently warm and collect some of the gas by downward displacement. How does the gas behave in contact with the air? Cautiously note the odor of the gas. Will it burn or support combustion? Is it soluble in water? Hold a piece of moist blue litmus paper in the gas. What change takes place? Dip litmus paper in the aqueous solution of the gas. Compare this solution with the hydrochloric acid of the laboratory. What is the composition of the latter? Can you show that hydrochloric acid contains hydrogen? (122)

Ex. **87.** Clean out the apparatus used in the last experiment and place some manganese dioxide in the flask (*A*). Add some strong hydrochloric acid and heat the mixture. Collect several portions of the gas by downward displacement, allowing as little of the gas as possible to escape into the room. What is the color of the gas? Cautiously note the odor. What is the name of the gas? Why was it so named? What was the source of this gas in the experiment? Will it dissolve in water? Place a colored flower and a moist piece of colored calico in a bottle full of the gas. Will chlorine bleach dry materials? To what is the bleaching due? Write the reaction between chlorine and water. What is formed by the action of chlorine on a metal? In what other ways may chlorides be formed? How may it be proved that hydrochloric acid is composed of hydrogen and chlorine? In what proportions by volume do the gases combine? In what proportion by weight?

Ex. **88.** In a test tube place some bleaching powder. Moisten with water and add sulphuric acid. What gas is given off? Write the reaction. Explain how bleaching powder is manufactured. Write the reaction. Of what use is bleaching powder?

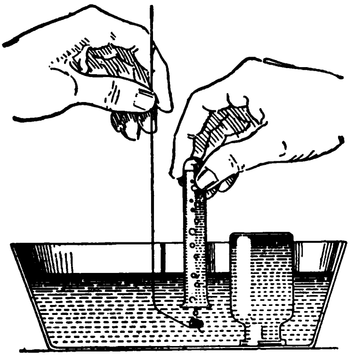


FIG. 98. — **The liberation of hydrogen from water by means of sodium.**

Ex. **89.** (Teacher.) Drop a piece of sodium the size of a grain of wheat on the surface of a pan of water. Do not stand too close as the sodium sometimes explodes. Touch the sodium while on the water with a taper. Is an inflammable gas evolved? Wrap a similar piece of sodium loosely in tin foil (to make it sink) and collect the gas in a test tube previously filled with water as shown in Fig. 98. Test the gas with a lighted splint. Write the reaction. Does the sodium liberate all the hydrogen of water? Is sodium chemically active? Why is it kept in coal oil? How is it prepared from common salt?

Ex. **90.** What compound remains when sodium acts upon water? What effect does the solution of sodium hydroxide have on litmus paper? Compare with limewater and hydrochloric acid. What is the common name for sodium hydroxide? For what is it used? What is the substance sold in cans under the name of potash? How is it manufactured?

Ex. **91.** Examine and describe ordinary washing soda. What is its chemical name and formula? What is meant by anhydrous sodium carbonate? Dissolve a little washing soda in water and test with litmus paper. What is a common use for this material? Make a saturated solution of sodium carbonate and pass carbon dioxide through it. (See Fig. 99.) What happens? What is the composition of the precipitate? Write the reaction. Add hydrochloric acid to washing soda and baking soda. Result? Which would yield the most carbon dioxide, a pound of washing soda or a pound of baking soda? Add baking soda to sour milk. What gas is given off? How can you prove that a gas is carbon dioxide?

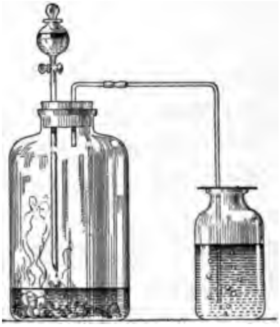


FIG. 99. — **Passing carbon dioxide through a solution of sodium carbonate.**

Ex. **92.** Dip a platinum wire in a solution of table salt and place it in the non-luminous flame. What happens? Hold a piece of glass rod in the flame. Have you any evidence that there is sodium in the glass? What is the test for sodium? Test a solution of common salt with the laboratory solution of silver nitrate. Is this a general test for chlorides? Ask the teacher to give you several substances and see if you can tell which are chlorides.

CHAPTER XV

ACIDS, BASES, AND SALTS

133. Acids. When sulphur trioxide was added to water (65), the solution had a sour taste and turned blue litmus paper red. The same thing was found to be true in a less marked way of sulphur dioxide (83) and of carbon dioxide (102); and the combinations made by these oxides with water were called *acids*. The oxides of several other elements behave in the same way as those of sulphur and carbon, and these compounds, known as acids, are of great importance. The acids mentioned must, from their method of preparation, contain hydrogen, oxygen, and some other element, but it has been shown that not all acids are formed by the union of an oxide with water. A few acids, one of which is hydrochloric acid, are formed by the union of one other element with hydrogen and contain no oxygen. All acids contain hydrogen, which may be replaced by a metal with the formation of a saltlike substance. Most acids will decompose carbonates, liberating the carbon dioxide with effervescence.

134. Bases. Not all oxides form acids when united with water. It has been seen that when calcium oxide or sodium oxide unites with water the compound formed has properties quite different from the acids. Solutions of these compounds turn red litmus paper blue, have a brackish taste, and do not decompose carbonates. These compounds are known as *bases* and contain a metal combined with oxygen and hydrogen.

135. Neutralization. A solution of sodium hydroxide, as has been said, turns red litmus paper blue, and a solution of hydrochloric acid turns blue litmus paper red. If, now, the hydrochloric acid solution is carefully added to the sodium hydroxide and the solution tested from time to time by placing a drop on litmus paper, a point can be found when the solution does not affect litmus paper at all. It will turn neither the red paper blue nor the blue paper red. The solution has neither a sour nor a brackish taste. In fact, all the characteristic properties of both acid and base have disappeared, and the two are said to have *neutralized* each other, the act being known as *neutralization*. When an acid and a base neutralize each other the action is *quantitative;* that is, it always takes exactly the same amount of the acid to neutralize a given quantity of the base. If the quantity of either acid or base is known, the other can easily be calculated.

136. Salts. If the solution resulting from neutralizing sodium hydroxide with hydrochloric acid is evaporated to dryness, it will be found that the white substance remaining in the dish is common salt (sodium chloride). If sulphuric acid is neutralized with sodium hydroxide, the product is sodium sulphate (Glauber's salt). Calcium hydroxide and sulphuric acid give calcium sulphate (gypsum). These compounds are examples of a class of substances known as salts. *When any acid is neutralized by a base one of the products formed is a salt; the other product in every case is water.* The formulas for sodium chloride ($NaCl$), sodium sulphate (Na_2SO_4), and calcium sulphate ($CaSO_4$) indicate that the formation of the salt really consists in the replacing of the hydrogen of the acid by a metal, such as sodium or calcium. The reaction for the formation of calcium sulphate is typical:

$$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2\ H_2O.$$

It will be seen that the metallic part of the base (Ca) replaces the hydrogen of the acid, and this hydrogen combines with the hydrogen and oxygen of the base to form water.

137. Definitions. *An acid* is a compound containing hydrogen, which may be replaced by a metal, the product formed being a salt.

A base is a compound which contains hydroxyl combined with a metal (or a basic radical) and which, when treated with an acid, easily exchanges its metal for hydrogen.

A salt is a compound formed when a metal replaces one or more of the hydrogen atoms of an acid. Salts may be formed also by the action of an acid (1) on a metal, and (2) on an oxide:

$$(1)\qquad Fe + H_2SO_4 \rightarrow FeSO_4 + H_2;$$
$$(2)\qquad FeO + H_2SO_4 \rightarrow FeSO_4 + H_2O.$$

138. Alkali. The bases or hydroxides formed by most of the metals are insoluble in water. The hydroxides of sodium, potassium, calcium, and a few others are soluble in water and show in a marked way the property of producing a blue color with certain vegetable dyes. They are also caustic in their action. These hydroxides are called *alkalies.* A substance that turns litmus or other of these vegetable dyes blue is said to have an *alkaline reaction.* A substance that turns the blue litmus red is said to have an *acid reaction,* and one that has no effect on litmus is said to be *neutral.* The organic coloring matter that is used to determine whether a substance is alkaline, acid, or neutral is called an *indicator*. Litmus is the most common indicator. Other important indicators are solutions of phenolphthalein and methyl orange — two coal tar compounds, and the aqueous extract of the cochineal insect.

139. Naming the Acids, Salts, and Bases. Those acids which are composed of only two elements (binary acids) are given the prefix *hydro* and the suffix *ic*. The salts of such acids have the suffix *ide*, and the prefix *hydro* is dropped; thus:

*Hydro*chlor*ic* acid (HCl). Sodium chlor*ide* ($NaCl$).

Most acids contain three elements — hydrogen, oxygen, and one other element — and are known as ternary acids. Such acids are given the ending *ic*, no prefix being used, and the salt is given the ending *ate*:

Sulphur*ic* acid (H_2SO_4). Calcium sulph*ate* ($CaSO_4$).

In case an element forms more than one acid containing oxygen, the one containing the larger percentage of oxygen is given the ending *ic* and the other the suffix *ous*. The salt in this latter case has the ending *ite*:

Sulphur*ous* acid (H_2SO_3). Sodium sulph*ite* (Na_2SO_3).

All the bases are called hydroxides, the name of the metal or basic radical being prefixed; for example, potassium hydroxide (KOH).

140. Normal, Acid, and Basic Salts. Some of the acids, as sulphuric acid (H_2SO_4) and carbonic acid (H_2CO_3), contain two atoms of hydrogen in the molecule. Sodium has a univalent atom which has the power of replacing only one hydrogen atom. It is possible to bring about a combination in which only one hydrogen atom of the acid is replaced, as for instance, $NaHSO_4$, in which half the hydrogen in sulphuric acid is replaced by sodium. Such a salt is known as an *acid salt* because it contains hydrogen, the characteristic constituent of acids. The salt in which all the hydrogen is replaced is a *normal salt.* For example, $NaHSO_4$ is acid sodium sulphate, sometimes called sodium hydrogen sulphate; and Na_2SO_4 is normal sodium sulphate. Baking soda, or sodium bicarbonate, $NaHCO_3$, is another example of an acid salt. The acid salts, when treated with a base, are converted into normal salts; and it has been shown that normal salts can be changed to the acid salts by the action of an acid:

$$NaHSO_4 + NaOH \rightarrow Na_2SO_4 + H_2O;$$
$$Na_2SO_4 + H_2SO_4 \rightarrow 2\ NaHSO_4.$$

(See also sodium bicarbonate and calcium bicarbonate.) The acids which have only one hydrogen atom in the molecule obviously cannot form acid salts. There are also a few salts in which the base is not completely neutralized by the acid. Such salts are known as *basic salts.*

141. Non-metals and Metals. It has been seen that certain elements form oxides which unite with water to produce acids, while the oxides of other elements form bases with water. The base-forming elements are for the most part metals (sodium, iron, zinc, copper). The acid forming elements, with a few exceptions, have no metallic properties. It is customary for the sake of convenience in discussion to divide the elements into two classes — *non-metals* and *metals.* Those elements that have at least one oxide that unites with water to form an acid are classified as non-metals. Those elements whose oxides form bases are known as metals. Unfortunately for the simplicity of this classification there are "border line" elements that are sometimes acidic and sometimes basic.

Note. An oxide which forms an acid when added to water is called an *anhydride.* Sulphur trioxide is sulphuric anhydride; that is, it is the oxide which forms sulphuric acid when added to water. Formerly the anhydrides were called acids, and even at the present time confusion is sometimes caused by the fact that in trade certain anhydrides are listed as acids.

The following elements are selected for discussion in this text:

Non-metals, or Acid-forming Elements	Metals, or Base-forming Elements
Hydrogen	Sodium
Oxygen	Potassium
Nitrogen	Calcium
Sulphur	Magnesium
Carbon	Copper
Chlorine	Silver
Phosphorus	Zinc
Silicon	Aluminum
Arsenic	Lead
Boron	Iron

Six of the non-metals and two of the metals have already been discussed. The remaining non-metals will next be considered, and afterward the metals and their salts.

Hydrogen really belongs in a class by itself but for convenience it is usually classed as a non-metal.

EXERCISES

Ex. **93.** Dissolve about two grams of sodium hydroxide in 100 cubic centimeters of water. Add 5 cubic centimeters of hydrochloric acid to 100 cubic centimeters of water. Test both solutions with red and blue litmus paper. Now carefully add the acid solution to the sodium hydroxide solution until the mixture will not change the color of either the red or the blue paper. Evaporate the mixture to dryness and determine whether the residue is common salt. What are the characteristics of the acids? What element do they all have in common? What is the characteristic group of the bases? How do bases and acids act on one another? What is meant by neutralization? When an acid neutralizes a base what is formed? How should you define an acid? A base? A salt? Mention three methods of forming salts.

Ex. **94.** What is meant by an alkali? By alkaline reaction? By acid reaction? When is a substance said to be neutral? What is an indicator? Try weak solutions of sodium hydroxide and hydrochloric acid with phenolphthalein indicator; with cochineal indicator; with methyl orange. Record the changes in color in each case. If red cabbage is available express some of the juice and determine whether it could be used as an indicator. Make a list of any substances at home that are acid or alkaline in reaction. What is the reaction of your perspiration? Of your saliva?

Ex. **95.** What is meant by a binary acid? How are these acids named? Give an example. When an element forms two acids how are they named? How are the salts named? How are the bases named? What is meant by an acid salt? Give an example. What is meant by an anhydride? Of the elements studied which are metals and which non-metals?

CHAPTER XVI

NITRIC ACID AND OXIDES OF NITROGEN

142. Chile Saltpeter. In northern Chile there are great beds of the substance commonly known as Chile saltpeter, or nitrate of soda. When pure, this material forms transparent, colorless crystals which are very soluble in water.

Fig. 100. — Mining nitrate of soda or Chile saltpeter.

In nature it is mixed with common salt and earth, from which it is separated by being dissolved in boiling water and then allowed to crystallize as the water cools. Nearly two million tons of nitrate of soda are exported from Chile each year. The larger part of this is used as a fertilizer; the remainder is utilized in the manufacture of nitric acid. If a particle of Chile saltpeter is placed in the flame of a Bunsen burner or an alcohol lamp the flame will be colored an intense yellow. This fact shows that this substance is a compound of sodium (131). It is evidently a salt, but since Chile saltpeter will not give the test for any of these acids (88, 101, and 132), the acid with which the sodium is combined is evidently different from any previously described. If Chile saltpeter is mixed with about its own weight of sulphuric acid and the mixture is gently warmed, a vapor escapes which, when condensed to a liquid, proves to be nitric acid, to which the formula HNO_3 has been assigned. Chile saltpeter, then, is the sodium salt of nitric acid; that is, it is sodium nitrate ($NaNO_3$).

143. Nitric acid is prepared commercially by heating sodium nitrate with sulphuric acid, upon which the following reaction takes place:

$$2\,NaNO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2\,HNO_3.$$

The apparatus in Fig. 101 is adapted to the preparation of the acid on a small scale. *A* is a small glass retort in which are placed 25 grams of sodium nitrate. About 15 cc. of sulphuric acid are added and the mixture is gently heated. The nitric acid distills over and is condensed in the test tube *B*, which is surrounded by cold water, preferably ice water.

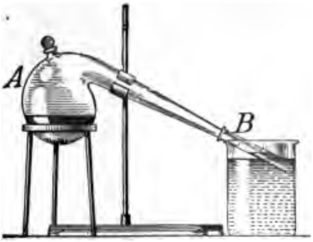


FIG. 101. — Apparatus used in making nitric acid.

Pure nitric acid is a colorless liquid about one and one half times as heavy as water. It gives off colorless fumes when exposed to the air. The concentrated nitric acid of the laboratory contains 68 per cent of HNO_3, the rest being water. In the above experiment the acid is slightly colored because, when nitric acid is boiled, a small part of it is decomposed according to the following equation:

$$2\,HNO_3 \rightarrow H_2O + 2\,NO_2 + O.$$

Nitrogen peroxide, NO_2, is a reddish brown gas which dissolves in the undecomposed nitric acid and colors it. The same decomposition takes place when nitric acid is exposed to strong light, and in consequence, the bottles in which concentrated nitric acid is stored often contain a reddish brown gas above the liquid, which is itself somewhat colored.

Strong nitric acid acts violently on many substances, especially those of animal and vegetable origin. It causes painful wounds when it comes in contact with the flesh, it eats through clothing, it burns wood and dissolves metals, and is, in fact, one of the most active of chemical substances. Even the dilute acid stains the skin and clothing yellow, and the stain cannot be removed. The greatest care should be exercised in working with nitric acid.

In dilute solutions, nitric acid has many of the characteristics of the other acids that have been studied. It has a decidedly sour taste, and it turns blue litmus paper red. It reacts with oxides, hydroxides, and carbonates to form salts, as, for example:

$$CuO + 2\,HNO_3 \rightarrow H_2O + Cu(NO_3)_2 \text{ (copper nitrate);}$$
$$NaOH + HNO_3 \rightarrow H_2O + NaNO_3 \text{ (sodium nitrate);}$$
$$CaCO_3 + 2\,HNO_3 \rightarrow H_2O + CO_2 + Ca(NO_3)_2 \text{ (calcium nitrate).}$$

When nitric acid acts on a metal, a salt is formed as would be expected, but no hydrogen is liberated; in this respect it differs from sulphuric and hydrochloric acids. Instead of hydrogen a gas which is one of the oxides of nitrogen is given off. The reaction with copper is usually written thus:

$$3\,Cu + 8\,HNO_3 \rightarrow 3\,Cu(NO_3)_2 + 4\,H_2O + 2\,NO.$$

To explain this action of nitric acid on metals it will be necessary to call attention to another property of nitric acid.

144. Nitric Acid an Oxidizing Agent. It was stated in Chapter VI that it is very difficult to make nitrogen combine with oxygen and hydrogen. It is also true that when this combination is brought about as in nitric acid the compound formed is very unstable. The formula for nitric acid (HNO_3) shows that it contains $\frac{48}{63}$, or over three fourths, of its own weight of oxygen. This oxygen is loosely held in the molecule and is freely given off to any readily oxidizable substance. This explains the action of strong nitric acid on animal and vegetable matter; for the changes caused in such materials by nitric acid are due to oxidation. Indeed, so readily does nitric acid give up its oxygen, that a piece of burning charcoal thrust beneath the surface of the strong acid will continue to burn, all the oxygen for the combustion being obtained from the acid.

When nitric acid acts on a metal, it is supposed that hydrogen is first liberated, as it is with other acids, but that the hydrogen is immediately oxidized to water by another portion of nitric acid which gives up part of its oxygen for that purpose. The first part of this reaction in the case of copper may be written as follows:

$$3\,Cu + 6\,HNO_3 \rightarrow 3\,Cu(NO_3)_2 + 6\,H.$$

The hydrogen thus evolved would then react with two more molecules of nitric acid, thus:

$$6\,H + 2\,HNO_3 \rightarrow 4\,H_2O + 2\,NO.$$

It will be seen that these two steps may be combined into the equation given in the preceding section.

145. Uses of Nitric Acid. Over 100,000 tons of nitric acid are used annually in the industries. It is used in the manufacture of nitroglycerin, which is the explosive constituent of dynamite. It is used also in making gun cotton, another explosive. It is required in the production of sulphuric acid by the chamber process (84) as well as in the manufacture of dyestuffs. The dilute acid is used in the refining of gold and silver and in the preparation of the copper plates from which etchings are printed.

146. Aqua Regia. A mixture of nitric and hydrochloric acid is called *aqua regia.* This expression means " royal water " and it was so named because it is the only acid that will dissolve the " noble metals," namely, gold and platinum. Its action depends upon the fact that the nitric acid oxidizes the hydrogen of the hydrochloric acid to water, and chlorine is liberated. The chlorine converts the metal into the chloride, which is soluble. In olden times all liquids were considered to be kinds of water. Nitric acid was called aqua fortis, or " strong water."

147. Nitrates. Nitric acid is monobasic; that is, the molecule contains one hydrogen atom. It forms salts with all the metals. All the nitrates are soluble in water. Most of them are colorless; copper nitrate, however, is blue and nickel nitrate is green. The nitrates are decomposed by heat. As a general rule the metal remains as the oxide while oxygen and an oxide of nitrogen are evolved, thus:

$$Cu(NO_3)_2 \rightarrow CuO + 2\,NO_2 + O.$$

In a few cases oxygen only is given off when the nitrate is heated, as for instance, in the case of sodium nitrate:

$$NaNO_3 \rightarrow NaNO_2 + O.$$

Owing to the ease with which they part with their oxygen, the nitrates, like the acid from which they are made, are good oxidizing agents.

148. Formation of Nitric Acid from the Air. Nearly all the nitric acid of commerce is manufactured from Chile saltpeter; but a small amount is now made by bringing about a direct combination of the nitrogen and the oxygen of the atmosphere. This is effected by means of powerful electric currents and is profitable only where water power makes very cheap electrical energy possible. At the present time the method is used to a limited extent in Norway. Most of the nitric acid so produced is treated with lime to form calcium nitrate, $Ca(NO_3)_2$, which is used as a fertilizer.

149. Test for Nitric Acid and Nitrates. To test for nitric acid or nitrates, dissolve the substance in water, place it in a test tube, and add a small quantity of a solution of copperas (ferrous sulphate). The tube is now inclined and some sulphuric acid is poured slowly down the side of the tube. The sulphuric acid, being heavier than the other liquids, will sink to the bottom of the tube without immediately mixing with the solution. If the substance being tested contains nitric acid or a nitrate, a dark ring (Fig. 102) will form at the point of contact between the sulphuric acid and the solution above it.

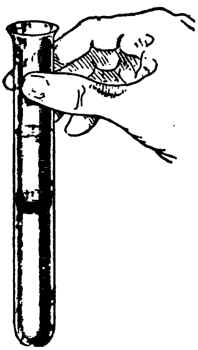


FIG. 102.—Testing for the presence of nitric acid.

150. Nitrites and Nitrous Acid. It was shown in Sec. 147 that when sodium nitrate is strongly heated one third of its oxygen escapes and the compound $NaNO_2$ is formed. This compound is sodium nitrite. It is more commonly prepared by melting sodium nitrate with lead. The lead extracts one third of the oxygen from the nitrate to form lead oxide, and sodium nitrite results, thus:

$$NaNO_3 + Pb \rightarrow NaNO_2 + PbO.$$

The nitrite may be dissolved in water, leaving behind the lead oxide, which is insoluble. Sodium nitrite is evidently the salt of an acid having the formula HNO_2, which should be named nitrous acid. This acid has never been prepared because it is so unstable that when it is liberated from its salts it immediately decomposes into various oxides of nitrogen. Most of the metals can be made to form nitrites, but none of them are of commercial importance except sodium nitrite, which is used in the manufacture of dyes. It is a solid that forms pale yellow crystals and is very soluble in water. The nitrates can be reduced to nitrites as noted above in the case of sodium nitrate. Likewise the nitrites can be easily oxidized to nitrates

$$NaNO_2 + O \rightarrow NaNO_3.$$

This reaction is of interest because nitrites are undoubtedly formed as intermediate compounds during the production of nitrates in nature, as will be shown in the next chapter.

151. Oxides of Nitrogen. Nitrogen forms five different combinations with oxygen. These five oxides of nitrogen are as follows:

Nitrous oxide	N_2O
Nitric oxide	NO or N_2O_2
Nitrogen trioxide	N_2O_3
Nitrogen peroxide	NO_2 or N_2O_4
Nitrogen pentoxide.............	N_2O_5

These oxides furnish one of the best examples of the law of multiple proportions.

152. Nitrous Oxide. Nitrous oxide (N_2O) is a colorless gas with a slightly sweetish taste. When inhaled it causes a slight intoxication which shows itself in the form of hysterical laughing. For this reason Sir Humphry Davy named it "laughing gas." Inhaled in larger quantities, it causes unconsciousness and is, therefore, used in certain minor surgical operations, particularly in extracting teeth. For this purpose it is condensed to a liquid and stored in steel cylinders. It is prepared by heating ammonium nitrate (NH_4NO_3), a substance which will be discussed in the next chapter. The equation is as follows:

$$NH_4NO_3 \rightarrow N_2O + 2\,H_2O.$$

153. Nitric Oxide and Nitrogen Peroxide. Nitric oxide (NO), as has been stated, is formed when nitric acid acts on some metals, as zinc or copper (143, 144). It is a colorless gas that is somewhat poisonous. Its most remarkable property is its power to combine directly with oxygen when the two are brought together. The reaction may be represented thus:

$$NO + O \rightarrow NO_2.$$

The product of this reaction is nitrogen peroxide (NO_2), which is a gas that has a reddish brown color and a disagreeable smell. It is very poisonous. If dilute nitric acid is poured upon some pieces of zinc or copper in a flask, the upper part of the vessel is filled with the colorless nitric oxide gas which, as it escapes into the air, takes up oxygen and changes to the reddish brown nitrogen peroxide. The latter will give up one half of its oxygen to any readily oxidizable substance and change back to the colorless nitric oxide. This power of the oxides of nitrogen to absorb oxygen from the atmosphere and transfer it to another substance is utilized in the manufacture of sulphuric acid by the chamber process (84).

EXERCISES

Ex. **96.** Examine crystals of sodium nitrate (Chile saltpeter). Has the substance the appearance of a salt? Test it in the flame. What metal is present? What is the source of the sodium nitrate of commerce? How is it purified? For what is it used?

Ex. **97.** In an apparatus as shown in Fig. 103 place a tablespoonful of sodium nitrate and sufficient sulphuric acid to cover the crystals. The bottle (*C*) should be filled with cold water or ice. Heat gently and examine the liquid that distills. Compare with the nitric acid of the laboratory. Is pure nitric acid colored? Why is the acid colored in this experiment? Why does nitric acid become colored when exposed to light? Write the reaction for the change. Dip a splinter of wood and a feather in nitric acid. How does nitric acid act on animal and vegetable tissues? On metals? How does its weak solution compare with solutions of other acids? How does it act on oxides? Hydroxides? Carbonates? Write the reactions.

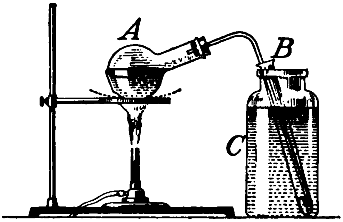


FIG. 103.—**Laboratory apparatus for making nitric acid.**

Ex. **98.** Heat nitric acid in a test tube with a piece of zinc or copper. Describe the result. What is the colored gas that is formed? Explain with equations the action of nitric acid on metals. Why is nitric acid said to be a good oxidizing agent? What uses are made of nitric acid? What is the mixture of nitric and hydrochloric acids called? What effect does it have on gold and platinum? What acid was formerly called aqua fortis? How is nitric acid made commercially? Can nitric acid be made from the nitrogen of the air? Where is this done?

Ex. **99**. Heat a nitrate in a hard glass test tube and prove that oxygen is evolved. Are the nitrates oxidizing agents? How can you test for the presence of a nitrate? Try the test with a little Chile saltpeter.

Ex. **100**. Heat sodium nitrate with lead. What compounds are formed? Write the reaction. What is the formula of nitrous acid? Has it ever been prepared in the pure state? Why? What change takes place in the nitrites when oxidized?

Ex. **101**. In a test tube heat a teaspoonful of ammonium nitrate. What is the gas which is given off? Give the formula. For what purposes is this gas used? Why is it called "laughing gas"?

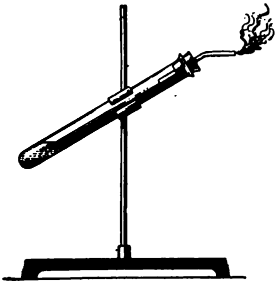


Fig. 104. — **Showing the formation of NO_2.**

Ex. **102**. Fit a test tube with a cork in which is fitted a small piece of glass tubing (Fig. 104). Place a little nitric acid and a piece of zinc in the test tube. Insert a cork and warm the acid. Is a gas evolved? What is the color of the gas in the test tube? What change takes place when it comes into contact with the air? Explain; give reaction. In what commercial process is this property of nitric oxide useful? How many oxides of nitrogen are known? Give the names and the formulas. How do they illustrate the law of multiple proportion?

CHAPTER XVII

AMMONIA AND ITS COMPOUNDS

154. Ammonia Water. One of the most familiar substances is the liquid so much used in the household under the name of ammonia water or spirits of hartshorn. This material (also called "aqua ammonia") consists of a gas dissolved in water, which like most other dissolved gases is almost completely driven off when the solution is heated. The gas may be prepared for examination in the apparatus shown in Fig. 105. The ammonia water is gently warmed in the flask *A* and, as some water may escape with the gas, it is passed through the tube *B*, which contains lumps of lime to absorb the water. As the gas is very soluble in water and is lighter than air, it is usually collected by upward displacement.

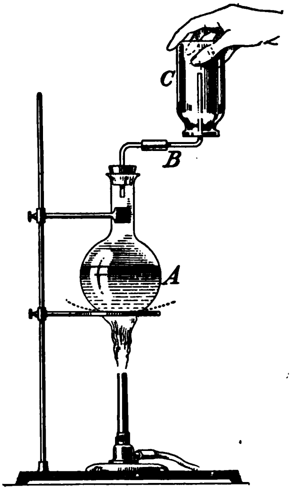


FIG. 105. — The production of ammonia.

155. Ammonia. The gas prepared in the above experiment is ammonia and is a compound having the formula NH_3. It is colorless and has an exceedingly pungent and penetrating odor. When inhaled it brings tears to the eyes, and in large quantities it may cause suffocation. It is about half as heavy as air and is very soluble in water, one volume of water dissolving 700 volumes of ammonia gas at ordinary temperatures. Ammonia is easily condensed to a liquid by pressure and cold. It will not burn in the air, nor will it support the combustion of a blazing stick; but in oxygen or in heated air it burns with a yellowish flame. A piece of moist red litmus paper is changed to blue if placed in ammonia gas.

156. Ammonia from Organic Matter. Whenever any animal or vegetable substance containing nitrogen is heated in a closed vessel so that the air does not have access to it, the nitrogen passes out of the compound as ammonia. This may be shown by heating bits of lean meat, horn, hoof, hair, peas, or beans in a hard glass test tube, as illustrated in Fig. 106. The escape of ammonia from the tube may be detected by the odor or by holding a piece of moist red litmus paper in the escaping gas. Ammonia was formerly made by the destructive distillation of the horns of the deer or hart and for that reason was termed spirits of hartshorn. Large quantities of ammonia are now produced as a by-product in the manufacture of animal charcoal from dried blood or bones. It should be noted that the nitrogen is not present in the animal or vegetable matter as ammonia, but the ammonia is formed during the decomposition of the organic matter which is brought about by the heat. When animal or

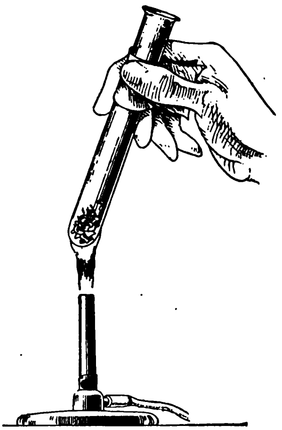


FIG. 106. — The production of ammonia from organic matter.

vegetable matter decays, the nitrogen present is liberated in combination with hydrogen as ammonia, and, consequently, the odor of ammonia is commonly noticed in stables, and in the vicinity of cesspools and manure piles.

157. Composition of Ammonia. It is very difficult to make nitrogen and hydrogen unite directly, although it can be done to a limited extent by passing electric sparks through a mixture of the two gases. Experiments with this method of production show that one volume of nitrogen always combines with three volumes of hydrogen, and, as nitrogen is fourteen times as heavy as hydrogen, the proportion by weight is 14 parts of nitrogen to 3 parts of hydrogen.

158. Manufacture of Ammonia. Bituminous coal contains small quantities of nitrogen (one to two per cent) and some hydrogen. When the coal is heated in the manufacture of illuminating gas, part of the nitrogen combines with hydrogen to form ammonia, which passes off with the gas. The illuminating gas is "washed" by being made to bubble through water, and the ammonia dissolves and remains behind in the so-called ammoniacal liquor. This liquor is then boiled with lime, and the ammonia is driven off and dissolved in pure water, forming the ammonia water of commerce. This is the most common method of manufacturing ammonia; and most of the household ammonia comes from this source.

159. Ammonia Combines with Water. When ammonia is absorbed by water, it is believed that the act is not one of mere solution, but that a chemical combination takes place between the ammonia and the water, thus:

$$NH_3 + H_2O \rightarrow NH_4OH.$$

The compound NH_4OH is named ammonium hydroxide, and the chemist regards ammonia water as a solution of ammonium hydroxide. This compound has never been separated because it is so unstable that any attempt to concentrate it and free it from water causes it to decompose into ammonia and water, thus:

$$NH_4OH \rightarrow NH_3 + H_2O.$$

This decomposition takes place at ordinary temperatures, and the odor of the solution is due to the escaping ammonia gas. Ammonia water may contain as high as 35 per cent by weight of ammonia gas, which, as has been shown, can be driven off by heat. A solution of this strength is lighter than water, the specific gravity being about 0.9. The solution ordinarily sold as household ammonia usually contains not more than 10 per cent of the gas.

Ammonia water is a strong alkali. It turns red litmus paper blue, and when rubbed between the finger and thumb has a slippery feel much like a weak solution of caustic soda. It is said to be a volatile alkali because it completely evaporates without leaving a residue; in which respect it differs from caustic soda, which is sometimes called a "fixed alkali." The fact that it leaves no residue gives ammonia water an advantage over the fixed alkalies for use in cleaning glassware and clothing. It is used also to soften water in the household and for other purposes where a milder alkali than caustic soda is required.

160. Ammonia Water Neutralizes Acids. If aqua ammonia is slowly added to a solution of hydrochloric acid, a point may be reached when the solution is neutral and has no effect upon litmus paper. If, now, this solution is evaporated to dryness, a white substance remains in the dish. This substance has much of the appearance and something of the taste of common salt. It is the substance known by the common name of sal ammoniac. Chemical analysis shows that its composition may be represented by the formula NH_4Cl, and the reaction that takes place may be expressed in the following equation:

$$NH_4OH + HCl \rightarrow NH_4Cl + H_2O.$$

This reaction is much like the one between sodium hydroxide and hydrochloric acid:

$$NaOH + HCl \rightarrow NaCl + H_2O.$$

Indeed, ammonia water can be used to neutralize any of the acids, and in each case a compound which closely resembles the corresponding compound formed from sodium hydroxide and the same acid is formed. The following compounds may be taken as examples:

	SODIUM HYDROXIDE	AMMONIUM HYDROXIDE
Forms with sulphuric acid	Na_2SO_4	$(NH_4)_2SO_4$
Forms with nitric acid	$NaNO_3$	NH_4NO_3
Forms with carbonic acid	Na_2CO_3	$(NH_4)_2CO_3$

These compounds are known as ammonium salts. It will be noted that where the symbol Na appears in the formulas of the sodium salts, the group NH_4 is found in the ammonium salts. A group like this, which acts as a unit in chemical reactions, is sometimes called a *radical.* The name ammonium has been given to the radical NH_4, and it is this radical which takes the place of the hydrogen of the acid when the salt is formed. In reality ammonium is an imaginary substance, for no one has succeeded in obtaining it by itself. There is good reason to believe, however, that it exists in ammonium salts and in ammonia water.

All ammonium salts are decomposed when treated with the fixed alkalies, and ammonia is given off. When ammonium chloride (sal ammoniac) is heated with calcium hydroxide, the following reaction takes place:

$$Ca(OH)_2 + 2\,NH_4Cl \rightarrow CaCl_2 + 2\,NH_3 + 2\,H_2O.$$

This reaction is quite commonly used in the preparation of ammonia for study in the laboratory.

161. Ammonium chloride or sal ammoniac (NH_4Cl) is prepared by passing ammonia gas from the "ammoniacal liquors" (158) into a solution of hydrochloric acid. It is a white, granular, or crystalline solid with a sharp, salty taste. The crude salt is sometimes called muriate of ammonia. It is used in certain kinds of electric batteries, in medicine, in soldering fluids, and in the textile industries. When heated, ammonium chloride is converted into a vapor without melting, and when the vapor comes in contact with a cold surface, it condenses in the form of minute crystals. This process of vaporizing and condensing a solid is called *sublimation,* and the solid is said to *sublimate.* All ammonium salts are either volatile or decompose when heated.

162. Ammonium nitrate (NH_4NO_3), made by passing ammonia into a solution of nitric acid, is a white crystalline solid, chiefly used in the preparation of nitrous oxide (152).

163. Ammonium sulphate ($(NH_4)_2SO_4$) is made by passing the ammonia of gas works into a solution of sulphuric acid.

$$2\,NH_4OH + H_2SO_4 \rightarrow (NH_4)_2SO_4 + 2\,H_2O.$$

It is a grayish yellow salt as produced commercially, and is the most widely used of all the ammonium salts. It is the starting point in the production of many of the ammonium compounds, and because of the fact that it is rich in nitrogen it is also largely used as a fertilizer.

164. Ammonium carbonate as found in commerce is an impure salt, being a mixture of acid ammonium carbonate (NH_4HCO_3) and a related compound. When fresh it is a transparent solid; but on exposure to the air it gives off ammonia and turns white. Smelling salts consist of lumps of ammonium carbonate covered with alcohol containing a little oil of lavender or other perfume. The commercial carbonate is sometimes used instead of baking powder. When it is heated the acid carbonate dissociates, forming water and the two gases — ammonia and carbon dioxide:

$$NH_4HCO_3 \rightarrow H_2O + NH_3 + CO_2.$$

It has an advantage over the baking powders in that it leaves no solid residue; but considerable experience is necessary to handle it successfully. It is used also in medicine and in scouring wool, and, in the household, for softening water. It is sometimes called crystal ammonia or solid ammonia, although the substance sold under the name of solid household ammonia is too often nothing but soda with a little ammonium carbonate added to give it the odor of ammonia. If no residue remains when a piece of the material is heated it is pure ammonium carbonate.

165. Ice Making with Ammonia. All liquids absorb heat when they evaporate (8). Ammonia absorbs very large quantities of heat in changing from the liquid to the gaseous form, and this fact is utilized in the manufacture of artificial ice (Fig. 107). Liquid ammonia is forced into a series of pipes (*A*) which are submerged in a large tank containing brine. The ammonia vaporizes and in so doing absorbs the required heat from the brine, which is thus cooled below the freezing point of pure water. If tin cans containing water are hung in this brine, the water will be frozen. As fast as the ammonia gas is formed in the cooling pipes it is removed by an exhaust pump and is liquefied by pressure and used over and over again. If there were no loss by leakage, the same amount of ammonia could be used indefinitely. When the ammonia condenses, it gives off heat, which is removed by running cold water over the condensing pipes (*B*). In cold storage warehouses (Fig. 108) the cold brine is made to circulate in iron pipes through the rooms to be cooled; or the pipes in which the ammonia is vaporized may be placed in these rooms instead of in the brine tank.

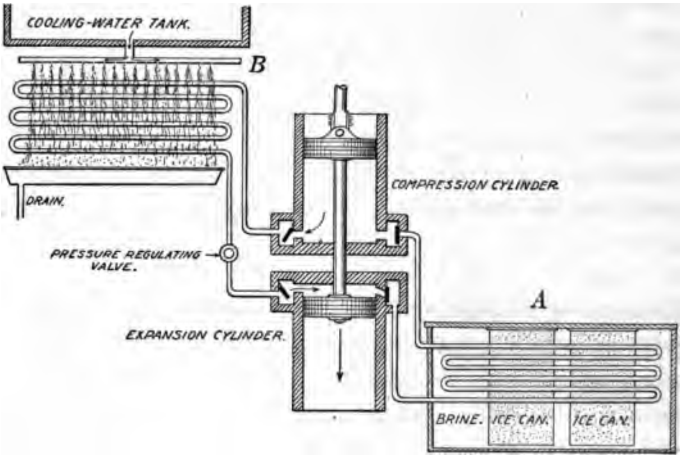


FIG. 107. — **The manufacture of artificial ice by the ammonia process.**

166. Occurrence of Ammonia. Although there are so many interesting and important uses for ammonia and its compounds, these substances are found only in very small quantities in nature. Ammonia is always found in minute traces in the atmosphere, because it is one of the products of the decay of plants and animals. It is dissolved in rain water and carried into the soil, where it is changed first into ammonium compounds and finally into nitric acid. It is found also in mere traces in most of the natural waters. Its presence in water in larger quantities is an indication that the water is contaminated by sewage. Owing to the instability of ammonium compounds and the ease with which they are changed to nitrates, there are no large deposits of ammonium salts as there are of nitrate of soda (142).

FIG. 108. — **Room in a cold storage warehouse.**

167. Test for Ammonium Salts. These salts are readily detected, since ammonia is evolved when they are treated with a dilute solution of caustic alkali, such as sodium hydroxide. The ammonia may be recognized by its odor or by the fact that it turns moist red litmus paper blue.

168. The Nitrogen Cycle. Although the worker in the laboratory has difficulty in making nitrogen unite with other elements, nature evidently has methods of bringing about this union. Small quantities of nitric acid are formed during electric storms, and this is carried into the soil by means of the rains. The combined nitrogen added to the soil in this way amounts to only three to eight pounds a year for each acre of ground. The principal factors in causing the formation of nitrogen compounds are the bacteria that live in the soil. An ounce of a good garden soil is said to contain at least one hundred fifty millions of bacteria. These are very small one-celled plants that can be seen only with the strongest microscope. Some of these bacteria have the power of bringing about the union of nitrogen, oxygen, and water to form nitric acid, which acid (as well as that in the rain water) usually unites with the calcium carbonate in the soil to form calcium nitrate, thus:

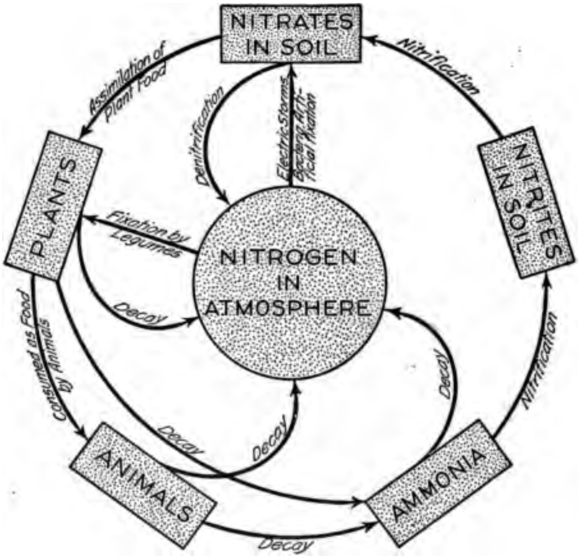


FIG. 109. — The nitrogen cycle in nature.

$$2\,HNO_3 + CaCO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2.$$

All plants need nitrogen in order to grow, and most of them obtain their nitrogen from the soil in the form of nitrates. The nitrogen compounds formed by these bacteria, therefore, are used by the higher plants. Some plants, however, have another way of obtaining the nitrogen they need. They belong to the family of plants known as legumes, which includes the clovers, alfalfa, peas, and beans. On the roots of these plants are found numbers of nodules or tubercles, which consist largely of masses of bacteria. These bacteria, while living on the roots of the legumes, have the power of causing the nitrogen of the air to form a chemical combination which the plant can utilize. Clovers and other legumes are frequently grown by farmers as a means of increasing the nitrogen compounds in the soil. This power of causing free nitrogen to enter into chemical combination is called fixation of nitrogen. Since animals get all the nitrogen in their bodies from the foods consumed, it follows that the bacteria are either directly or indirectly responsible for nearly all the nitrogen compounds found in nature.

Bacteria are responsible also for other changes in nitrogen compounds. When the plants or animals die, their bodies decay, as do also the waste products of the animal body. The decay is caused by other kinds of bacteria and results in the breaking down of the complex nitrogen compounds found in plants and animals. Sometimes the nitrogen is liberated in the form of pure nitrogen or of ammonia and finds its way into the atmosphere. This change is called *denitrification.* If the decay takes place in the soil it is more likely to result in the formation of nitrates, which can again be utilized by growing plants. This change from complex nitrogen compounds to nitric acid and nitrates is termed *nitrification.* The change is caused by at least three kinds of bacteria and takes place in three steps:

(1) the formation of ammonium compounds from the organic nitrogen compounds; (2) the change of ammonium compounds into nitrous acid (HNO_2), which unites with a base in the soil to form a nitrite—probably calcium nitrite, $Ca(NO_2)_2$; (3) the oxidation of the nitrite to a nitrate,

$$Ca(NO_2)_2 + 2\,O \rightarrow Ca(NO_3)_2.$$

It will thus be observed that nitrogen passes through a cycle (Fig. 109) much like that described for carbon, although the amount of nitrogen involved in these changes is much smaller than the carbon of the carbon cycle.

169. Another Method of Fixing Nitrogen. When calcium carbide (CaC_2) is strongly heated in a current of nitrogen, a substance is formed which has the formula $CaCN_2$:

$$CaC_2 + N_2 \rightarrow CaCN_2 + C.$$

This new substance is variously named, *nitro-lime, lime nitrogen,* and *calcium cyanamide.* It is made on a commercial scale by means of the electric furnace. It is a hard, gray-black mass resembling coke and, as the formula shows, is rich in nitrogen. It is used as a fertilizer, but before the nitrogen can be utilized by plants it must be oxidized to nitrate, a process which readily takes place in the soil (624).

EXERCISES

Ex. **103.** Perform the experiment described in paragraph **154.** Why is the gas collected by upward displacement? How does the gas affect the eyes and nose? Test the gas with a piece of red litmus paper. Will the gas burn or support combustion? Does it dissolve in water? Can it be condensed to a liquid? What is its name and formula?

Ex. **104.** Heat bits of meat, hair, horn, and some beans in hard glass test tubes. Test the vapor from the tubes with moist red litmus paper. What is the source of the ammonia? Was the nitrogen present in the above materials as ammonia? Is ammonia formed by the decay of animal and vegetable matter? What is the composition by weight of ammonia? How is the ammonia water of commerce manufactured?

Ex. **105**. Examine a sample of ammonia water from home. How does it feel when rubbed between the fingers? What is the odor? The reaction with litmus paper? Is it an alkali? Why is it called a volatile alkali? Is its volatility any advantage in cleaning? Does ammonia gas merely dissolve in water or form a chemical compound with it? Write the reaction. Is this compound stable? Why does ammonia water always give the odor of ammonia gas? How much ammonia does the ordinary household ammonia water contain?

Ex. **106**. To a dilute solution of hydrochloric acid, add dilute ammonia water until the mixture is neutral or faintly alkaline. Evaporate to dryness. What remains? Give the formula. Write reaction between ammonium hydroxide and hydrochloric acid. Will ammonia water form salts with other acids? Compare with sodium salts of the same acids. What is a radical? What is the name of the radical NH_4? Does this radical exist? Mix a little ammonium chloride with lime and heat in a test tube. Try the same experiment with ammonium sulphate and sodium hydroxide. How do ammonium salts act when heated with the fixed alkalies?

Ex. **107**. Place a teaspoonful of ammonium chloride in a long test tube. Hold the test tube in an inclined position and heat. What happens to the chloride? Does it condense again at the top of the test tube? What is meant by sublimation? What are the commercial uses of ammonium chloride?

Ex. **108**. Examine some ammonium sulphate. How is it made commercially? What are its uses? Does ammonia combine with carbonic acid? Examine crystals of crude ammonium carbonate. Why is it called sal volatile? What happens to it when heated? Write the reaction. What uses are made of it? Examine a commercial sample of "solid household ammonia" and determine whether it contains soda. How should you test for ammonium salts?

Ex. **109**. Explain by help of a diagram how liquid ammonia is utilized in ice making.

Ex. **110**. Where and to what extent is ammonia found in nature? Why is it not more abundant? Discuss the cycle of nitrogen in nature. What is formed when calcium carbide is strongly heated in a current of nitrogen? What use is made of calcium cyanamide?

CHAPTER XVIII

PHOSPHORUS, PHOSPHORIC ACID, ARSENIC

170. PHOSPHORUS has already been used in the experiments with oxygen (37). It is an element that is never found in the free state, and is usually combined with calcium and oxygen in the form of calcium phosphate.

Phosphorus is slightly yellow and translucent. It can be cut like wax at ordinary temperatures, and it melts at 44° C. It is insoluble in water but dissolves freely in carbon bisulphide. It is very poisonous; and in factories where phosphorus is made, the workmen are frequently poisoned by it.

FIG. 110.—Sticks of phosphorus.

In contact with the air phosphorus gives off fumes which emit light visible in a dark room. This phenomenon suggested the name phosphorus, which is derived from the Greek and means "light bearer." Although other substances act in the same way, this property was first observed in connection with phosphorus and the phenomenon is, therefore, called *phosphoresence.*

The streak of light left by a match when rubbed on any surface in a dark room is due to the phosphorus in the match. Phosphorus takes fire when rubbed or cut; hence it must be handled with great care. It is kept under water, and should be cut under water and never held in the hand, since the heat of the hand is sufficient to ignite it. The burns caused by phosphorus are very difficult to heal. In a finely divided state phosphorus ignites spontaneously. If a little phosphorus is dissolved in carbon bisulphide and the solution poured on a piece of filter paper, the phosphorus will take fire as soon as the carbon bisulphide has evaporated.

171. Red Phosphorus. When ordinary phosphorus is exposed to the light for a long time, it becomes opaque and darker in color, and finally dark red. The change to red phosphorus can be hastened if the yellow variety is heated in a sealed tube to about 250° C. Red phosphorus is strikingly different in its behavior from the yellow. It is not very active. It does not change in the air, and must be heated to a comparatively high temperature before it will combine with oxygen. It is insoluble in carbon bisulphide and is not poisonous. Red phosphorus may be changed back to the ordinary variety by heating it to 300° C. in an atmosphere of nitrogen.

172. Preparation of Phosphorus. The ash produced by burning bones is the source of most of the phosphorus of commerce. Phosphorus cannot be readily prepared in the laboratory. Commercially it is made by heating bone ash with sand and charcoal. A complicated reaction takes place during which phosphorus is liberated as the element and escapes from the retort as a vapor which is condensed under water. It is purified by redistillation and cast into sticks under water. These sticks are usually about half an inch in diameter and 7.5 inches long.

173. Matches. The principal use of phosphorus is in the manufacture of matches. The ordinary parlor matches are made by dipping small pieces of wood into melted paraffin and then into a mixture of phosphorus, manganese dioxide, and glue. By rubbing such matches on a rough surface enough heat is generated by the friction to cause the phosphorus to ignite. This sets fire to the paraffin, which in turn kindles the wood. The manufacture of these matches is prohibited in some countries because they are so liable to take fire and because the workmen in the factories are so often poisoned by the phosphorus. Safety matches, or Swedish matches, contain no yellow phosphorus. The head of this kind is a mixture of potassium chlorate, antimony sulphide, and glue. The side of the box is coated with red phosphorus, glue, and powdered glass. When the head of the match is drawn over this prepared surface, a little of the phosphorus is torn off, catches fire, and ignites the match. Safety matches cannot easily be ignited except on the prepared surface, although they can be lighted by drawing them rapidly over glass.

174. Phosphorus burns with the formation of a dense white cloud, which condenses into a white, snowlike solid. Either variety of phosphorus burns readily, but the combustion of the yellow is much more violent. The product of combustion, which is the same in either case, has the composition P_2O_5 and is called phosphorus pentoxide. This oxide of phosphorus is very deliquescent, quickly drawing moisture from the air. It combines vigorously with water, with a hissing sound. It is often used in the laboratory to dry gases.

175. Phosphoric Acid. When phosphorus pentoxide is added to hot water, the solution has a sour taste and turns blue litmus paper red. This is due to the fact that the phosphorus pentoxide has combined with the water to form phosphoric acid, H_3PO_4:

$$P_2O_5 + 3\,H_2O \rightarrow 2\,H_3PO_4.$$

Phosphoric acid, also called orthophosphoric acid, when pure is a white solid that is very soluble in water. The commercial article contains a little water and is a thick, sirupy liquid somewhat resembling pure sulphuric acid in appearance and weight. While it can be prepared as indicated above, it is usually made by the common method for preparing acids; namely, by the action of sulphuric acid on one of the salts of phosphoric acid. The salt ordinarily used is calcium phosphate, $Ca_3(PO_4)_2$, which is a constituent of bones and of the phosphate rocks.

$$Ca_3(PO_4)_2 + 3\ H_2SO_4 \rightarrow 3\ CaSO_4 + 2\ H_3PO_4.$$

The calcium sulphate is insoluble and is filtered off.

176. Salts of Phosphoric Acid. The formula for phosphoric acid, H_3PO_4, shows that each molecule contains three hydrogen atoms. From what was said about normal and acid salts (140) it is evident that one, two, or three of these hydrogen atoms might be replaced by a metal. There are, therefore, three possible phosphates of each metal. The three phosphates of sodium, for example, with their formulas and chemical names are as follows:

Mono-sodium phosphate, NaH_2PO;
Di-sodium phosphate, Na_2HPO_4;
Normal or Tri-sodium phosphate, Na_3PO_4.

The second salt, di-sodium phosphate, is the most common, and is the one used in the laboratory and in medicine. It is usually known simply as sodium phosphate. It crystallizes with twelve molecules of water of crystallization:

$$Na_2HPO_4 \cdot 12\ H_2O.$$

In the case of a metal like calcium, which has a valence of two, the manner in which the base and acid combine to form the salts is not so apparent. For the bivalent calcium just to neutralize phosphoric acid with its three hydrogen atoms, it must be assumed that three atoms of calcium react with two molecules of phosphoric acid, thus:

$$3\,Ca + 2\,H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6\,H.$$

The three calcium salts and their formulas are,

Monocalcium phosphate, $CaH_4(PO_4)_2$,
Dicalcium phosphate, $Ca_2H_2(PO_4)_2$,
Normal or Tricalcium phosphate, $Ca_3(PO_4)_2$.

177. Occurrence of Phosphorus. The bulk of all the phosphorus found in nature exists in the form of tricalcium phosphate. The bones of animals are 80 per cent tricalcium phosphate. It occurs in large deposits in a mineral known as apatite, and in a more impure form it is found in the phosphate rocks of Florida, Tennessee, South Carolina, Arkansas, Kentucky, Idaho, Utah, Wyoming, and Montana. The rocks from which the soils were formed contained some tricalcium phosphate also, and consequently it is present in small quantities in all soils.


FIG. 111. — Hydraulic mining of phosphate rock in Florida.

Phosphorus is found in plant and animal tissue. The plants derive it from the phosphate in the soil and build it up into complex organic compounds. Animals eat the plants and get their phosphorus in that way. When plant residues and the animal bodies and manures decay, the phosphorus is returned to the soil, where it is again oxidized to phosphates; thus it completes the phosphorus cycle.

Some of the iron ores contain phosphorus, which is objectionable in the manufacture of steel. To remove it lime is added, and the phosphorus remains with the calcium in the slag, which is ground to a fine powder and sold as *basic slag,* Thomas phosphate, or odorless phosphate (236).

178. Fertilizers. Phosphorus in the form of a phosphate is absolutely essential to plant growth. Very few soils contain sufficient phosphorus for a maximum crop, and as a large part of the phosphorus in the plant is stored in the seeds, which are removed from the land, the amount of phosphorus in the soil is being constantly decreased. A part of the phosphorus is returned to the soil in animal manures, but never a sufficient amount to restore that removed by the crops. To maintain a satisfactory crop yield, the farmer must add some form of calcium phosphate to the soil. This is sometimes done by the use of bones, which are ground to a fine powder (bone meal), or by the use of basic slag. The phosphate rocks are also ground to a fine powder called *floats* and applied directly to the land or mixed with the animal manures. Tricalcium phosphate, however, is very insoluble in water, and to make it more available it is customary to treat it with sufficient sulphuric acid to bring about the following reaction:

$$Ca_3(PO_4)_2 + 2\ H_2SO_4 \rightarrow 2\ CaSO_4 + CaH_4(PO_4)_2.$$

Monocalcium phosphate, $CaH_4(PO_4)_2$, is soluble in water and can, therefore, be much better distributed in the soil than the insoluble natural phosphate. The calcium sulphate (gypsum) and monocalcium phosphate produced by the above reaction are not separated, but the mixture is dried and ground, and sold under the various names of *superphosphate, acid phosphate,* or *acidulated rock* (610).

179. The term phosphoric acid as used in fertilizers does not mean H_3PO_4 but the anhydride P_2O_5 (141). Formerly calcium phosphate was considered to be a combination of lime and phosphoric anhydride. The formula was written $3\,CaO \cdot P_2O_5$. The name phosphoric acid was then given to the oxide P_2O_5, and although this use of the name has been discontinued by chemists, it has persisted in trade. When a fertilizer, then, is said to contain 14 per cent of phosphoric acid, it means that it contains calcium phosphate equivalent to 14 per cent of P_2O_5.

180. Test for Phosphoric Acid or Phosphate. If the phosphate is insoluble in water, it should be dissolved in dilute nitric acid. To the solution to be tested nitric acid should be added unless it was used to dissolve the substance. Add about 2 cc. of the ammonium molybdate test solution and warm gently. If phosphoric acid is present, a yellow precipitate will be formed. Since the composition of the precipitate depends upon the temperature, however, no formula can be assigned to it.

181. Arsenic. The name arsenic, which is sometimes used for the white material sold in the drug stores, should be confined to the element arsenic, which is a steel gray, brittle solid, with a metallic appearance. It is found sparingly in the uncombined state, but its compounds with sulphur and with the metals are very abundant. It is not in common use, but it is sometimes added to lead as a hardener in the manufacture of shot.

182. White Arsenic. When arsenic, or any of its ores, is burned, the arsenic is converted into the oxide:

$$2\,As + 3\,O \rightarrow As_2O_3.$$

Arsenic trioxide (As_2O_3) is the substance ordinarily called *arsenic*, or *white arsenic*. It is found in trade as a white powder which has no odor and a faintly sweet taste. It is a deadly poison and should be handled with great care. It is slightly soluble in water and the solution is poisonous to both plants and animals. Arsenic compounds are used almost entirely for the destruction of vermin and insect pests.

183. Arsenites. When arsenic trioxide is boiled with sodium hydroxide, the following reaction takes place:

$$As_2O_3 + 6\,NaOH \rightarrow 2\,Na_3AsO_3 + 3\,H_2O.$$

Sodium arsenite, Na_3AsO_3, is soluble in water and is used as a basis for the preparation of other compounds of arsenic. As it is very soluble it is more intensely poisonous than white arsenic. The corresponding calcium salt, calcium arsenite, $Ca_3(AsO_3)_2$, is insoluble and is often used as a spray to kill potato bugs and other insects. It is made by adding a solution of sodium arsenite to calcium hydroxide (slaked lime).

Sodium arsenite upon oxidation is changed to the arsenate, Na_3AsO_4. Evidently this compound is the sodium salt of arsenic acid, H_3AsO_4, while sodium arsenite is the salt of arsenious acid, H_3AsO_3. Neither of these acids is of any importance; but several of their salts are in common use. The arsenites and arsenates of all metals except sodium and potassium are insoluble. Some of them will be studied in connection with the metals from which they are formed.

EXERCISES

Ex. 111. What is the appearance of ordinary phosphorus? Is it ever found in nature in the elemental condition? Is it soluble in water? What effect does it have upon the men who work with it? What happens to phosphorus when exposed to the air? Why was it named phosphorus? Why is it stored under water? When ordinary phosphorus is heated in a sealed tube what change takes place? Is the red phosphorus very active? In what other respects does it differ from ordinary phosphorus? How is phosphorus prepared commercially? How are the safety matches made?

(*Note.* All experiments with yellow phosphorus should be performed by the teacher.)

Ex. **112.** Burn a little red phosphorus in a dry wide-mouth bottle (Fig. 112). What is formed? What is the composition of the white fumes? The name? Pour a little water into the bottle and shake. Does the white material dissolve? Test solution with blue litmus paper. Write equation for action of phosphorus pentoxide on water. Examine the phosphoric acid of the laboratory. How is it prepared commercially? How many atoms of hydrogen in the molecule? How many sodium phosphates are possible? Names and formulas?

FIG. 112. — **Burning phosphorus in bottle.**

Ex. **113.** Dissolve a little sodium phosphate in water. Heat and add a few drops of the ammonium molybdate reagent of the laboratory. What happens? This is the test for a phosphate. Mix some rock phosphate with water. Filter and test the filtrate for phosphates. Moisten another sample of the rock phosphate with sulphuric acid. After ten minutes add water, stir, and filter. Test this filtrate for phosphates. Have you any evidence that the sulphuric acid made the rock phosphate soluble? Why is rock phosphate treated with sulphuric acid in making fertilizers? What is meant by acid phosphate, or acidulated rock? Discuss the occurrence of phosphorus in nature. What is meant by the term phosphoric acid as used in the fertilizer trade?

Ex. **114.** What is the composition of the substance known as white arsenic? What is its chemical name? What is formed when it is boiled with sodium hydroxide? What is formed when the arsenites are oxidized? What practical use is made of calcium arsenite?

CHAPTER XIX

SAND, SILICON, BORAX

184. Sand. The term sand is sometimes used to designate any gritty material consisting of small angular fragments of rocks or minerals. In a more restricted sense sand consists of small particles of more or less pure silica. In its pure form silica crystallizes in beautiful six-sided prisms and is called quartz, or, sometimes, rock crystal. These crystals are often so clear that they can be used for making spectacle lenses or as substitutes for the diamond. Sea sand is often almost exclusively fragments of quartz. White sands are practically pure silica, while in the yellow sands, or the variously tinted sandstones, the silica is colored by iron oxide or some other metallic oxide.

FIG. 113. — Quartz crystal.

Chemically, silica is an oxide of the element silicon and is called silicon dioxide, which has the composition represented by the formula SiO_2. Quartz is the most common of minerals and constitutes 18 per cent of the crust of the earth. In the form of the mineral, *quartzite*, it forms many mountains, and the sandstones also consist almost entirely of silica. Several of the valuable stones and precious gems, as onyx, or carnelian, agate, jasper, flint, amethyst, rhinestone, and opal, consist of silicon dioxide. Infusorial earth consists of the skeletons of minute aquatic organisms and is nearly pure silicon dioxide. It is employed as a scouring and polishing material, and is used to absorb nitroglycerin in the manufacture of dynamite. Silica also occurs in the leaves and stalks of grasses, cereals, and bamboos and other canes. The plant known as equisetum (horsetail) contains so much silica that it is often used for scouring and is called scouring rush. Silica constitutes about 40 per cent of the ash of the feathers of birds and is found in the hair of animals.

185. Silicon. The element silicon is never found in the free state, although, next to oxygen, it is the most abundant element. The solid crust of the earth contains 28 per cent of silicon; for it is found in all varieties of granite, sandstone, gneiss, clay, and shale. It may be prepared by heating the oxide (quartz or white sand) with powdered magnesium.

$$SiO_2 + 2\,Mg \rightarrow Si + 2\,MgO.$$

Silicon made in this way is an amorphous brown powder insoluble in water and in all the common acids. It is used to some extent in the steel industry as a reducing agent. When heated to a high temperature, it burns and forms silicon dioxide (SiO_2).

186. Water Glass. When clean white sand is melted with sodium carbonate, carbon dioxide escapes, and a compound remains which is called sodium silicate (Na_2SiO_3).

$$Na_2CO_3 + SiO_2 \rightarrow Na_2SiO_3 + CO_2.$$

Sodium silicate is soluble in water and is known as water glass. In commerce it is found as a thick, sirupy solution. It is employed in fireproofing cloth and wood, as a cement, as a " filler " in laundry soaps (462), and in preserving eggs. For the latter purpose the commercial solution is diluted with nine times its own volume of water, and the eggs are kept immersed in the liquid until used. If the eggs are perfectly fresh at the beginning, they can be preserved for several months in this solution.

187. Silicic Acid. If sulphuric or hydrochloric acid is added to a concentrated solution of water glass, silicic acid separates in the form of a jelly, thus:

$$Na_2SiO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2SiO_3.$$

Silicic acid has never been prepared in a pure state because when evaporated to dryness it decomposes, and pure white sand (silicon dioxide) remains in the dish:

$$H_2SiO_3 \rightarrow H_2O + SiO_2.$$

188. Glass. When limestone is heated at a high temperature with sand, calcium silicate is formed and carbon dioxide is given off:

$$CaCO_3 + SiO_2 \rightarrow CaSiO_3 + CO_2.$$

If both calcium and sodium carbonates are melted with sand, the product is a mixture of sodium and calcium silicates which, upon cooling, forms ordinary window glass. The ingredients are used in about the following proportions: clean sand, 150 pounds; soda, 50 pounds; limestone, 25 pounds. Window glass is made by blowing a lump of the glass into a hollow cylinder, which is then cut lengthwise and allowed to spread open upon a

FIG. 114. — Glass blowing.

flat surface. Plate glass has the same composition but is made by pouring the molten glass upon a large table and rolling it with a hot iron roller, and subsequently grinding and polishing it.

When potash is used in place of the soda, a very hard glass is formed known as Bohemian glass, which is much used for chemical apparatus. Lamp chimneys, lenses, and cut glass are made of flint glass. This is a silicate of lead and potassium, made by melting potassium carbonate and lead oxide with sand. Glass is colored by adding different substances which dissolve in the molten mass. The green color of common glass bottles is due to the iron in the impure sand used; copper and cobalt produce different shades of blue; manganese dioxide gives a pink or violet color, and certain copper compounds and gold make the glass ruby red.

189. Natural Silicates. The salts of silicic acid are called silicates. They make up a large part of the earth's crust, the silicates of aluminum, calcium, potassium, sodium, magnesium, and iron being the most abundant. Nearly all the common rocks, with the exception of limestone and dolomite, are silicates, as well as many of the minerals and a few of the precious gems. Mica, clay, slate, asbestos, soapstone, feldspar, meerschaum, garnet, emerald, topaz, and beryl are all silicates. Some of these silicates are salts of the silicic acid (H_2SiO_3) mentioned above, but many of them are evidently salts of acids which are very much more complex. Silicates are known, for instance, which correspond to acids having the formulas H_4SiO_4, $H_6Si_2O_7$, $H_4Si_3O_8$, and many others. None of these more complex acids have ever been isolated, but their salts are well known, and some of them will be described when the metals from which they are formed are studied.

When the silicic acid contains four or more hydrogen atoms, it is quite common to find that two or more metals have replaced the hydrogen atoms to form a mixed salt. For example, feldspar is a silicate of potassium and aluminum ($KAlSi_3O_8$); and mica is another silicate of potassium and aluminum ($KAlSiO_4$).

190. Decomposing the Silicate. All the silicates, with the exception of those of sodium and potassium, are insoluble in water. Most of them are also quite insoluble or very slightly soluble in acids. If an insoluble silicate is mixed with sodium carbonate and the mixture is heated, it melts or fuses, and sodium silicate and an insoluble carbonate are formed; that is, the sodium and the other metals change places.

$$CaSiO_3 + Na_2CO_3 \rightarrow CaCO_3 + Na_2SiO_3.$$

The sodium silicate can be dissolved in hot water, leaving the metals of the original silicate behind as carbonates, which can be dissolved in hydrochloric acid. This is the method used in the laboratory to decompose the silicates for purposes of analysis.

191. Test for Silica or Silicates. If the material is insoluble in water or acids, it is fused with sodium carbonate and then treated with boiling water as described in the previous paragraph (190). Hydrochloric acid is then added to the hot water solution with the result that the gelatinous silicic acid separates; or the acidulated solution may be evaporated to dryness, upon which a white residue of silica (SiO_2) remains which cannot be dissolved in water and hydrochloric acid. No substance but silicic acid behaves in this way when treated with hydrochloric acid.

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presented, and the detection of the basic or metallic part of the salt will be studied in Chapter XXVIII.

199. Examination of a Solid. If the substance under examination is an element, it is probably carbon, or sulphur, as these are the only familiar non-metallic elements that are solids. *Carbon may be recognized by its black color* (89–94) and by the fact that carbon dioxide is formed when it burns. Ignite the substance and hold over it a glass rod which has been dipped in limewater (101). *Sulphur may be recognized by its yellow color* (59). Verify by dissolving in carbon bisulphide and recrystallizing (60).

If the material is not one of the above elements, place a small quantity in a test tube, moisten it with sulphuric acid, warm it gently, and note the result.

(*a*) A colorless gas may be given off. The odor of sulphur dioxide indicates that the substance is a sulphite (86 and 88). The odor of hydrogen sulphide indicates a sulphide (87). If the gas is odorless, it is probably carbon dioxide from a carbonate. Test it with a glass rod that has been dipped in limewater (101). A gas may be given off which is colorless but which fumes when breathed upon. This indicates hydrochloric acid from a chloride (121). Verify by dissolving a small portion of the original substance in water and adding a drop of nitric acid and a few drops of silver nitrate (132). It may be nitric acid from a nitrate. Verify by adding a bit of zinc and heating (153), or by testing a water solution of the original substance with ferrous sulphate and sulphuric acid (149).

(*b*) A yellow gas which does not fume may be chlorine from bleaching powder (124); in which case the gas will readily bleach litmus paper or a bit of colored cloth or a flower (123).

(*c*) No gas evolved indicates a sulphate (88), a phosphate (180), a borate (196), a silicate (191), or a basic oxide.

(*d*) The substance may be an ammonium salt. Test by heating a small portion with a solution of sodium hydroxide and noting the odor (167).

EXERCISES

Ex. **121.** Obtain samples of single unknown substances from the teacher and test them carefully according to the plan outlined in this chapter. Read the chapter carefully before beginning the experiment. Look up all cross references. The substances may be any of the following:

sulphur
a sulphide
a sulphite
a sulphate
a chloride
carbon
a carbonate
a nitrate
a phosphate
a borate
an ammonium salt

Make a careful record of the results of each test.

(*Note.* The teacher should make this chapter the basis of a thorough review of the preceding chapters.)

CHAPTER XXI

POTASSIUM

200. The metal potassium resembles sodium in most of its properties. It is a soft solid, which will float on water. The freshly cut surface has a silvery-white metallic luster. It acts upon water even more energetically than does sodium, causing the evolution of hydrogen and forming potassium hydroxide:

$$K + H_2O \rightarrow KOH + H.$$

The heat evolved by this reaction is so great that it ignites the liberated hydrogen. Like sodium it must be stored in coal oil to prevent its absorbing moisture and oxygen from the air.

201. Occurrence of Potassium. The metal is never found free, but its compounds are widely distributed. Many of the rocks from which soils are formed contain potassium compounds, and consequently potassium is present in small quantities in all soils. Potassium is one of the essential constituents of plant food and is always found in plants. When vegetable material is burned, the potassium remains in the ashes as potassium carbonate (K_2CO_3). Formerly wood ashes were the principal source of potassium. The impure potassium carbonate dissolved from the wood ashes is called potash. Some of the giant seaweeds that grow along the Pacific coast contain as much as 35 per cent of their dry weight of potassium chloride. Much potassium is obtained from the potash deposits of Europe. These deposits are made up of sixteen or more different salts, and the beds, which are nearly 3000 feet thick, were probably formed by the evaporation of sea water.

The element potassium is prepared by electrolysis of its compounds in the manner described under sodium (125).

FIG. 115. — Mining potash salts.

202. Potassium Hydroxide. Potassium hydroxide (KOH) is a white, brittle substance resembling sodium hydroxide, and is prepared by the same methods (126). It absorbs moisture when exposed to the air and is used for removing both water and carbon dioxide from gases. In chemical behavior it is like sodium hydroxide, and as the latter is much cheaper than potassium hydroxide it is more commonly used commercially. The common name for potassium hydroxide is *caustic potash.*

203. Potassium Chloride. Potassium chloride (KCl) resembles the corresponding sodium compound in appearance and chemical behavior. The chloride is found in large quantities in the European deposits and is used as the starting point in the production of most of the other potassium compounds. The crude salt is sold as a fertilizer under the trade name of *muriate of potash.*

204. Potassium Carbonate. Potassium carbonate (K_2CO_3) is commonly prepared from wood ashes. The ashes are placed in a barrel or other receptacle, and water is poured on them and drawn off at the bottom. The lye thus obtained contains potassium carbonate, or *potash,* and is often used in making soft soap. The refined potassium carbonate is called *pearl-ash.* It has been stated that the substance sold as potash is very often sodium hydroxide (126).

FIG. 116. — Leaching potash from wood ashes.

205. Potassium Sulphate. Potassium sulphate (K_2SO_4) occurs in combination with other salts of potassium in the European deposits. It is sold as a fertilizer under the name of *sulphate of potash.* It is also used in medicine and in preparing ordinary alum.

206. Potassium Nitrate. Potassium nitrate (KNO_3) is commonly called *saltpeter* or, sometimes, *niter.* When organic matter containing nitrogen decays in the presence of bases, nitrates are formed (168). Advantage was formerly taken of this fact to produce saltpeter artificially. Refuse animal matter was mixed with earth and wood ashes and the pile was moistened with liquid manure from the stable. After two or three years the nitrate that had developed in the pile was dissolved out with water and purified. In some hot, dry climates saltpeter is formed in the soil near the villages in quantities sufficient to be extracted on a commercial scale. At the present time, however, most of the potassium nitrate is made by treating a hot solution of sodium nitrate with potassium chloride.

$$NaNO_3 + KCl \rightarrow KNO_3 + NaCl.$$

Potassium nitrate is a white solid which occurs in long slender crystals. It gives off oxygen readily when heated and is, therefore, a good oxidizing agent. It is used to some extent in medicine and in the preservation of meat, but its principal use is in the manufacture of black gunpowder.

207. Gunpowder is a mixture of potassium nitrate (75 per cent), charcoal (15 per cent), and sulphur (10 per cent). The ingredients are moistened with water and thoroughly mixed by grinding, and the mixture is then dried. When gunpowder burns in a closed space, a large amount of gas is suddenly formed. One gram of powder yields 280 cc. of gas, and the heat produced causes a great expansion of the gas. The reaction is approximately as follows:

$$2\,KNO_3 + 3\,C + S \rightarrow 3\,CO_2 + 2\,N + K_2S.$$

The explosion is due to the suddenness with which the gases are generated and the heat is developed. Smokeless powder (314) has now replaced black gunpowder in warfare.

208. Potassium Chlorate. Potassium chlorate ($KClO_3$) was used in the preparation of oxygen (28). At high temperatures it decomposes into oxygen and potassium chloride:

$$KClO_3 \rightarrow KCl + 3\ O.$$

Potassium chlorate forms flat white crystals and tastes like saltpeter. It is used to prepare oxygen and in the manufacture of fireworks and matches. In the form of chlorate of potash tablets it is used as a remedy for sore throat. It is prepared by the electrolysis of a hot solution of potassium chloride:

$$KCl + 3\ H_2O \rightarrow KClO_3 + 6\ H.$$

209. Potassium Cyanide. Potassium cyanide (KCN) is a white solid which smells like bitter almonds. It is extremely poisonous. It is used in photography and in extracting gold from its ores. When acted upon by sulphuric acid, it yields a gas having the formula HCN:

$$2\ KCN + H_2SO_4 \rightarrow K_2SO_4 + 2\ HCN.$$

This gas (HCN) is called hydrocyanic acid. Its solution in water is known as prussic acid. Hydrocyanic acid is a deadly poison and is sometimes used in fumigating trees to kill scale insects and also in the fumigation of greenhouses. It must be used with extreme care.

210. Test for Potassium Compounds. A beautiful violet color imparted to the Bunsen flame is the test for potassium compounds. This test is easily applied to the pure compounds, but if sodium is present the yellow sodium flame hides the violet color of the potassium. If the flame is viewed through a piece of blue (cobalt) glass or through a thin layer of indigo solution, the violet color can readily be seen, while the yellow color is not transmitted.

EXERCISES

Ex. **122.** What substance previously studied does potassium resemble? How does it react with water? Is it more or less energetic than sodium? Discuss the occurrence of potassium in nature. What is potash? What compound of potassium is found in wood ashes? What is the principal source of potassium compounds at the present time? Try the flame test on a potassium salt. What is the result?

Ex. **123.** What is caustic potash? Allow a small piece of potassium hydroxide to remain exposed to the air on a watch glass. What happens? What substance with which you are familiar does potassium chloride resemble in appearance and chemical behavior? Where is crude potassium chloride obtained? What use is made of it chemically? By what common name is it generally known?

Ex. **124.** Place some wood ashes on a filter and pour on a little water. How does the liquid which runs through the filter feel when rubbed between the fingers? Evaporate the filtrate. What is the composition of the residue? Have you ever seen soft soap made with lye from wood ashes? Do wood ashes which have been exposed to the weather contain much potash?

Ex. **125.** What is the chemical composition of saltpeter? The chemical name and formula? How was it formerly produced? How is it produced at the present time? For what is it used? How is black gunpowder, or blasting powder, made? What causes the explosion when gunpowder is ignited? What is the probable reaction?

Ex. **126.** What is the formula of potassium chlorate? What changes does it undergo when heated at a high temperature? Have ready half a teaspoonful of potassium chlorate and a like quantity of manganese dioxide on separate papers. Place the chlorate in a dry test tube and heat cautiously until the chlorate is melted. Test with a glowing splint for oxygen. If not heated too strongly, no oxygen will be evolved. Remove the flame and at once drop the manganese dioxide into the melted chlorate and test for oxygen. Why is the manganese dioxide used with potassium chlorate in making oxygen? Does the manganese dioxide undergo any change?

CHAPTER XXII

MAGNESIUM AND ZINC

MAGNESIUM

211. Occurrence of Magnesium. Magnesium is widely distributed and ranks close to calcium in amount. The stone known as *dolomite,* or magnesian limestone ($CaCO_3 \cdot MgCO_3$) is well known. Some mountain ranges are largely of dolomite and beds of it cover thousands of square miles in the Mississippi Valley. Magnesite ($MgCO_3$) occurs frequently. Magnesium carbonate is usually found to some extent in all limestones. Magnesium is found in sea water, in many mineral waters, and in several of the European potash salts. It is a component of serpentine, talc soapstone, asbestos, meerschaum, and some other silicates.

212. Magnesium is a lustrous, silvery-white metal having a specific gravity of only 1.75. It is produced from its compounds by electrolysis. It is ductile, and when hot may be drawn into wire or ribbon, the latter being a common commercial form. It burns with a dazzling white light. Magnesium powder is an ingredient of photographic flashlight powders, in which the magnesium is mixed with about twice its own weight of powdered potassium chlorate.

213. Magnesium oxide (MgO) is the white bulky powder formed when magnesium burns. It is usually manufactured by heating the carbonate, just as lime is made from limestone. It is often called *magnesia,* or *calcined magnesia.*

It is used in medicine for certain forms of dyspepsia and as an antidote for poisoning by mineral acids. It does not form the hydroxide as readily as calcium oxide does, and the lime made from magnesium limestone is not so desirable for building purposes as that from calcium limestone.

214. Magnesium sulphate ($MgSO_4 \cdot 7\,H_2O$) is the salt known as *Epsom salts.* It was first recognized in the mineral spring at Epsom, England. It is used in medicine as a purgative; also as a coating for cotton cloth, and in dyeing.

215. Magnesium chloride ($MgCl_2$) is even more deliquescent than calcium chloride. When exposed to the air, it soon absorbs sufficient moisture to dissolve itself. It is present in small quantities in all natural deposits of common salt, and the tendency of table salt to become moist and cake is due to the presence of magnesium chloride. A small amount of sodium bicarbonate added to salt will prevent caking. Starch is sometimes used for the same purpose.

216. Test for Magnesium Compounds. Make a small cavity in a piece of charcoal and fill it with the substance to be tested. Moisten with water and heat strongly in the blowpipe flame (Fig. 125). Cool, add a drop of cobalt nitrate solution, and heat again. Cool and examine. Magnesium compounds leave a pink or pale flesh-colored residue in the charcoal.

ZINC

217. Preparation of Zinc. Zinc is found as zinc blende (ZnS) and smithsonite ($ZnCO_3$). From the carbonate ore, zinc oxide (ZnO) is obtained by heating. From zinc blende, the oxide is produced by roasting the ore which removes the sulphur and leaves the oxide:

$$ZnS + 3\,O \rightarrow ZnO + SO_2.$$

The oxide is then reduced by heating with powdered coal:

$$ZnO + C \rightarrow Zn + CO.$$

This method of extracting zinc from its ores should be carefully considered. The ores of most metals consist of the carbonates, oxides, or sulphides; and the method of handling them is in general the same as that described for zinc. In the case of the other metals only the forms of the furnaces and other details vary. The art of extracting metals from their ores is called *metallurgy.* The metallurgy of zinc involves the roasting of the ores to produce the oxide, which is then reduced with carbon.

218. Pure zinc (Zn) is a bluish white metal with a specific gravity of 7.1. It can be rolled into thin malleable sheets. When melted and poured into water, it forms thin flakes and in this condition is called *mossy zinc.* Sheet zinc is used as a lining for tanks and sinks and to protect floors beneath stoves. In sticks or plates it is used in electric batteries. Iron dipped into melted zinc becomes coated with zinc and is called *galvanized iron.* The zinc protects the iron and prevents rusting. About two thirds of the zinc produced is used in this way. In the laboratory it is used to prepare hydrogen. It is used also in the manufacture of brass, German silver, and other alloys (255).

219. Zinc oxide (ZnO) is a white powder obtained by roasting the carbonate, or by burning the metal. It is commonly known as zinc white. It is used as a pigment in white paints and has the advantage over white lead of not darkening from exposure to hydrogen sulphide, as zinc sulphide is white, while lead sulphide is black. Zinc white, however, has only three fourths the covering power of white lead. Zinc oxide is used as a filler in the rubber of automobile tires. It is also a constituent of zinc ointment.

220. Other Zinc Compounds. Zinc chloride ($ZnCl_2$) is a white deliquescent solid. The aqueous solution is used for cleaning metal surfaces before soldering. The largest use of the chloride is in wood preservation. Zinc sulphate ($ZnSO_4 \cdot 7\ H_2O$), commonly known as white vitriol, is used in medicine and in the dyeing and printing of cloth.

221. Testing for Zinc Compounds. (1) Zinc is the only common metal that forms a white sulphide that is insoluble in water. If the substance is not soluble in water, dissolve in hydrochloric acid, neutralize the acid with ammonia water, and add hydrogen sulphide. A white precipitate indicates zinc. (2) Fill a small cavity in a piece of charcoal with the substance. Moisten it with water and heat it strongly in the blowpipe flame. Cool it and moisten it with a drop of cobalt nitrate solution, then heat it again. Cool and examine. Zinc compounds leave a green incrustation.

EXERCISES

Ex. **127**. Ignite a short piece of magnesium ribbon and note how it burns. What is the white compound formed? What use is made of metallic magnesium? Discuss the occurrence of magnesium in nature. What is the most abundant compound of magnesium? What is calcined magnesia? Is magnesium lime as valuable for building as the calcium lime? What is the common name for magnesium sulphate? What is the effect of the presence of a small quantity of magnesium chloride in table salt? What is meant by the statement that magnesium chloride is very deliquescent?

Ex. **128**. Test a sample of a magnesium compound (216). Result?

Ex. **129**. Give the properties of zinc. For what is it used? How does it react with dilute acid? Write reaction for preparation of hydrogen from zinc and sulphuric acid. How is zinc separated from its ores? Give reactions in the preparation of zinc from the sulphide. How is zinc oxide prepared? How used? Test a sample of zinc oxide according to paragraph 221. Name articles at home that contain zinc.

CHAPTER XXIII

ALUMINUM

222. ALUMINUM is a bluish white metal with a specific gravity of 2.6, which is about one third that of iron. It is ductile and malleable and can be readily drawn into a wire or pressed into thin sheets. Aluminum does not change in the air, and this property, combined with its low specific gravity, makes it useful in the manufacture of articles in which lightness is important. Its attractive appearance has led to its extensive use as an ornamental metal. It is extensively used for the manufacture of cooking utensils. The powdered metal is used in making aluminum paint. It is a good conductor of electricity and is coming into use in electric work as a substitute for copper, especially in long-distance wires. A small quantity of aluminum added to cast

FIG. 117. — **Lightweight camp cooking outfit of aluminum.**

iron prevents the formation of bubbles and air holes. The metal is trivalent and readily displaces hydrogen from hydrochloric acid:

$$Al + 3\,HCl \rightarrow AlCl_3 + 3\,H.$$

Nitric acid and dilute sulphuric acid act upon it very slowly. Concentrated sulphuric acid dissolves it, forming the sulphate and sulphur dioxide. The metal also dissolves in strong alkalies. Acids and alkalies should be avoided in using cooking vessels made of aluminum.

223. Occurrence. Aluminum forms 7 per cent of the earth's crust, and next to oxygen and silicon it is the most abundant element. As the silicate it is found in feldspar, mica, and clay. As the oxide (Al_2O_3), it occurs in corundum and emery, which are used as abrasives. The sapphire, the ruby, the oriental topaz, and the amethyst are aluminum oxide colored by traces of impurities. Cryolite, a fluoride of aluminum and sodium (Na_3AlF_6), is an important compound. Aluminum occurs also as the hydrated oxide called bauxite ($Al_2O_3 \cdot 3\,H_2O$).

224. Preparation. Although aluminum is very abundant in nature, it is never found in the free state. The cost of the earlier methods of separating it from its ores was so great that until recent years it was almost a curiosity. Since aluminum is a stronger reducing agent than carbon, the metal cannot be prepared by the method used for zinc. It was formerly made by heating

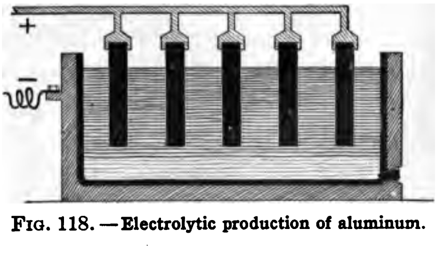

FIG. 118. — Electrolytic production of aluminum.

the chloride with metallic potassium. The method now used (Fig. 118) was discovered in 1886 by Charles W. Hall, an American youth, just out of college and only twenty-two years of age. Hall's method consists in passing an electric current through a mass of molten cryolite to which bauxite has been added. Under these conditions the bauxite is decomposed into aluminum and oxygen. This process has reduced the price of the metal from $90 a pound in 1886 to about twenty cents a pound at present. If a cheap method could be discovered to prepare it from ordinary clay, the metal could be put to many new uses.

225. Aluminum sulphate ($Al_2(SO_4)_3 \cdot 18\,H_2O$) is a white crystalline solid, readily soluble in water. It is used in purifying water and sewage, as a mordant for fixing dye-stuffs on fabrics, and as a sizing material in the manufacture of paper.

226. Ordinary Alum. When solutions of aluminum sulphate and potassium sulphate are mixed and evaporated, transparent, colorless, glassy crystals are formed. This solid is alum and has the composition represented by the formula $KAl(SO_4)_2\ 12\,H_2O$. Alum is very soluble in water, and the solution has an acid reaction and a sweetish, puckery taste. When heated it loses its water of crystallization and some sulphuric acid and falls into a white powder or porous mass known as burnt alum, which is used in medicine.

227. Other Alums. It will be noted that alum is composed of a univalent metal (potassium) and a trivalent metal (aluminum) combined with the sulphuric acid radical. Any univalent atom may be substituted for potassium, and any trivalent atom (as the iron or chromium atom) may take the place of aluminum. No matter what the combination *may* be, the crystalline form, the water of crystallization, the acid reaction, and the puckery taste are the same. The following are the better known alums:

$KAl(SO_4)_2 \cdot 12\ H_2O$ — potassium alum;
$NH_4Al(SO_4)_2 \cdot 12\ H_2O$ — ammonium alum;
$NH_4Fe(SO_4)_2 \cdot 12\ H_2O$ — iron alum;
$KCr(SO_4)_2 \cdot 12\ H_2O$ — chrome alum.

The aluminum alums are used in the dyeing industry, in the manufacture of paper, and, improperly, in baking powder (426). They are also used in fireproofing wood and cloth. Chrome alum is used in the tanning industry and as a hardener in the fixing bath used in photography.

228. Clay, Pottery, Porcelain. Clay is an impure aluminum silicate formed by the action of carbon dioxide and water on the feldspars, during the weathering of granite rocks. The products of the decomposition are chiefly an insoluble aluminum silicate and a soluble alkaline silicate. The latter is largely washed away. The aluminum silicate is pure clay or kaolin ($H_4Al_2Si_2O_4$). *Pure kaolin is a white, powdery mass.* Ordinary clay contains many impurities, such as quartz and compounds of iron, calcium, and magnesium.

FIG. 119. — The interior of a pottery kiln.

All kinds of clay form a stiff plastic mass which can be molded into any shape. When dried it shrinks considerably, and when strongly heated it shrinks still further and forms an infusible mass which is not attacked by water or acids, and which can no longer be made into a paste with water. In this way bricks, pottery, and porcelain are made. The red color in bricks and common pottery is due to iron in the clay.

FIG. 120. — **Brick making in the old way, India.**

As burned clay is very porous, pottery is generally glazed by throwing salt into the oven in which pottery is being fired. The steam from the clay decomposes the salt, giving NaOH and HCl, and the alkali combines with some of the clay, forming a fusible silicate, which melts and covers the pottery or brick, and on cooling becomes a hard glassy film.

For porcelain pure kaolin mixed with feldspar is used. After the porcelain is fired, it is glazed by being covered with a thin cream of powdered feldspar and water and heated to a white heat. The feldspar melts and penetrates the porcelain, forming a thoroughly adherent glaze.

229. Ultramarine is a deep blue material used as a paint pigment, for laundry blue, in making blue tinted paper, and in correcting the yellow shade of linen, starch, sugar, and paper stock. It is made by heating together kaolin, sodium carbonate, sulphur, and charcoal. Formerly it was prepared by powdering the blue mineral, lapis lazuli.

230. Test for Aluminum. When an aluminum compound is strongly heated in a blowpipe flame, and the resulting white mass is moistened with a drop of cobalt nitrate solution and again heated, it becomes sky blue.

EXERCISES

Ex. 130. Give the general properties of aluminum. Does it rust or tarnish? How does it compare with iron in weight? For what is it used? Why should acids and alkalies be avoided in using cooking vessels made of aluminum? How does aluminum occur in nature? How is the metal prepared? Why can it not be reduced with carbon as in the case of zinc? What effect did Hall's discovery have on the price of aluminum? What aluminum articles can you find at home?

Ex. 131. Dissolve about ten grams of aluminum sulphate in the least possible amount of boiling water. Dissolve 3 grams of potassium sulphate in the same way. Mix the clear, hot, saturated solutions in a shallow dish, and allow the mixture to cool undisturbed. Remove and examine the best crystals which form. Are they aluminum sulphate? Potassium sulphate? Has a new compound been formed? What is this substance? Test to prove that these crystals contain potassium, aluminum, and the sulphate radical. Give the formula for ordinary alum. Test a solution of alum with blue litmus paper. Give names and formulas of three other alums. Mention some uses of the alums.

Ex. 132. What is clay and how formed? What is the appearance of pure kaolin? Why are some clays colored? Work some clay into a plastic mass with water. Heat a piece of this mass to a high temperature. How has the heat affected the plasticity of the clay? How are bricks and common pottery made? What is meant by glazing the pottery? From what is porcelain made? How is it glazed?

Ex. 133. What is ultramarine and how is it made? For what is it used? How was it formerly prepared? Test a crystal of alum for aluminum according to paragraph 230. What is the result?

CHAPTER XXIV

IRON

231. IRON (Fe) is undoubtedly the most useful of all the metals. It rarely occurs in the free state, but as it is easily prepared from its ores it has been known from the early ages. The principal ores of iron are the oxides, hematite (Fe_2O_3) and magnetite (Fe_3O_4), and the carbonate ($FeCO_3$). Iron in the metallic state has been found in meteors.

FIG. 121. — **Mining iron ore, Minnesota.**

232. Extraction of Iron from Its Ores. The chemical process of obtaining iron from its ores is a very simple one. Whatever ore is used, it is first converted into the oxide by roasting, if it is not already in the form of the oxide (217). It is then reduced in a blast furnace by carbon in the form of coke or hard coal. The carbon removes the oxygen from the oxide. The blast furnace (Fig. 122) is from 25 to 90 feet high and from 15 to 18 feet wide in the widest part. Alternate layers of ore and fuel are introduced at the top of the furnace. Since many ores of iron contain earthy impurities, limestone is always placed in the furnace with the iron and coke. This combines with the earthy materials to form slag, which is somewhat similar to glass. The molten iron, being the heavier liquid, sinks to the bottom and is drawn off and cast into bars about three feet long and six inches thick, which are known as pigs. The slag is sometimes used in the manufacture of Portland cement.

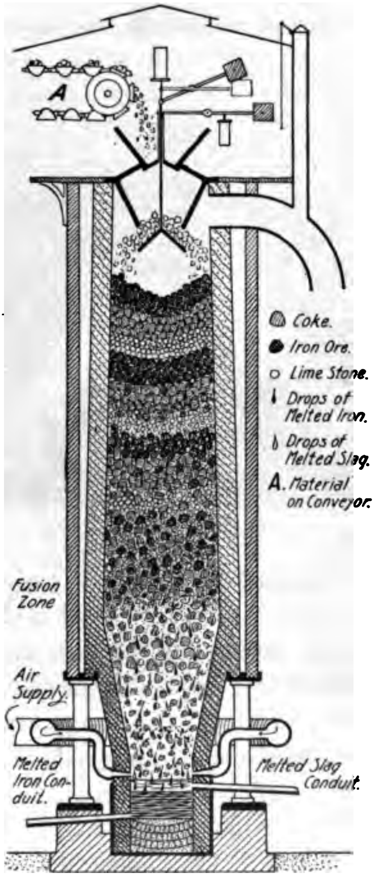

FIG. 122. — A blast furnace.

233. Cast Iron. The pig iron drawn off from the furnace is always impure, containing some phosphorus, silicon, sulphur, and carbon. It is brittle and easily melted, and is used for casting, being known as *cast iron*.

234. Wrought Iron. When practically all of the impurities are removed from iron, it is no longer brittle. It becomes malleable and at a red or white heat it can be hammered or pressed into any desired shape. It is now known as wrought *iron.* It can also be *welded;* that is, two pieces of the metal can be united when hammered or rolled together.

235. Steel is iron that contains from 0.8 to 2.5 per cent of carbon. Steel can be forged like wrought iron. A very important property of steel is its power of being tempered, or rendered hard or soft at will. When heated to redness and suddenly plunged into cold water, it is rendered very hard and brittle. If heated and slowly cooled, it is made soft, and by regulating the temperature at which it is tempered, almost any degree of hardness, toughness, or elasticity may be obtained.


FIG. 123. — Casting pig iron.

236. Basic Slag. In the manufacture of steel from iron containing phosphorus the converter is lined with dolomite (211), which absorbs the phosphoric oxide produced during the process, and forms a basic calcium phosphate. When the lining has absorbed all the phosphorus it can take up, it is removed, pulverized, and sold as a fertilizer under the name of *basic slag,* or *Thomas phosphate.* Phosphorus makes steel brittle; hence the necessity for removing it.

237. Rusting of Iron. All kinds of iron oxidize readily in moist air, even at ordinary temperatures. To protect the iron from the air and moisture and thus from rusting, it is covered by a coat of paint, or it is galvanized (218), tinned, or nickel-plated.

238. Iron Has Two Valences. One atom of iron may hold two or three atoms of a univalent element in combination. In other words, iron may be bivalent or trivalent, according to chemical conditions. Thus there are two chlorides, $FeCl_2$ and $FeCl_3$; two nitrates, $Fe(NO_3)_2$ and $Fe(NO_3)_3$; two sulphates, $FeSO_4$ and $Fe_2(SO_4)_3$. The compounds in which iron appears to be bivalent are called ferrous compounds; for example, ferrous chloride and ferrous nitrate. The compounds of trivalent iron are called ferric compounds; for example, ferric chloride and ferric sulphate.

239. Change of Ferrous to Ferric Compounds. Ferrous compounds are changed to ferric compounds by contact with air, and oxidizing agents, such as nitric acid. When, for instance, ferrous hydroxide ($Fe(OH)_2$) is exposed to air while suspended in water, it slowly changes to ferric hydroxide ($Fe(OH)_3$):

$$2\,Fe(OH)_2 + H_2O + O \rightarrow 2\,Fe(OH)_3.$$

So, also, when ferrous chloride is left standing in a solution of hydrochloric acid, it changes to ferric chloride, and the change is rapidly effected by boiling with a little nitric acid, which gives up oxygen:

$$2\,FeCl_2 + 2\,HCl + O \rightarrow 2\,FeCl_3 + H_2O.$$

240. Ferrous sulphate ($FeSO_4 \cdot 7\,H_2O$) is the compound commonly known as *copperas* or *green vitriol.* It is formed by the action of sulphuric acid on iron. It crystallizes in pale green crystals. It is used as a purifier of water, as a disinfectant, as a reagent for killing weeds, in the dyeing industry, and in the manufacture of writing ink.

241. Inks. The common black writing inks are made by treating a solution of ferrous sulphate with a solution of tannic acid obtained from nutgalls. A little gum arabic is usually added, and a preservative to prevent the ink from molding.

242. Sulphides of Iron. When iron and sulphur are heated together, they unite to form ferrous sulphide (FeS). It is a black substance which is used in the laboratory to prepare hydrogen sulphide. A sulphide having the formula FeS_2 is abundantly found in nature and is known as iron pyrites. It is a yellow crystallized substance sometimes called *fool's gold.* When strongly heated, the sulphur is oxidized to sulphur dioxide, and the iron is left in the form of the oxide. Iron pyrites is commonly used as a source of sulphur in the manufacture of sulphuric acid (84).

243. Iron compounds are very widely distributed in nature. All soils contain small quantities of iron. The yellow and red color of clays are due to the presence of iron, as are also the colors of many sandstones. Iron is found in plant and animal tissues in minute traces. The formation of the green coloring matter of plants (chlorophyll) is dependent upon the presence of iron. The blood of animals contains traces of iron. In fact, neither plants nor animals can live without iron, although the amount needed by them is exceedingly small.

244. Test for Iron. Tannic acid or an infusion of nutgalls forms a blue-black color and a very finely divided precipitate, which remains suspended in the liquid.

EXERCISES

Ex. **134.** What are the more common ores of iron? What can be said about the usefulness of iron? Does it ever occur in the free state? Is its use by man of recent or ancient origin? How is iron extracted from its ores? What is pig iron? Cast iron? What is the principal difference between cast iron and wrought iron? What is meant by welding? What is the purest common form of iron? In what marked way does steel differ from wrought iron? Visit a blacksmith shop and ask the smith to illustrate the tempering of steel.

Ex. **135.** Why is the presence of phosphorus in steel objectionable? How is the phosphorus removed in making steel? What is basic slag? Dissolve a little basic slag in nitric acid and test for phosphorus with molybdate reagent.

Ex. **136.** What causes the rusting of iron? Will iron rust if perfectly dry? How may it be protected to prevent rusting? What compounds does iron make with sulphur? What is fool's gold?

Ex. **137.** Dissolve a small amount of iron in hydrochloric acid. Divide the solution into two parts and to one add ammonium hydroxide until strongly alkaline. Filter. Explain the change in color of the precipitate when exposed to the air. Write the reaction. To the second part of the above solution add a little hydrogen peroxide. Explain the effect of the peroxide on the solution. Add ammonium hydroxide in excess as above. How does the precipitate compare with that from the first solution? What two classes of compounds does iron form?

Ex. **138.** What is the chemical name and formula of copperas? How is it made? For what is it used? Add a little tannic acid to a solution of copperas. What happens? What is the test for iron? Discuss the distribution of iron compounds in nature. Is iron necessary to plant and animal life? How many articles can you find at home that are made of iron?

CHAPTER XXV

LEAD

245. Occurrence and Metallurgy of Lead. Lead is found in nature principally as the sulphide (PbS), known as galena, or galenite, a heavy black mineral which crystallizes in cubes having a bright metallic luster. In the process of obtaining lead from the ore, the ore is first roasted in order that a part of the sulphide may be converted into the oxide and sulphate:


FIG. 124. — Galena.

$$PbS + 3\,O \rightarrow PbO + SO_2,$$
$$PbS + 4\,O \rightarrow PbSO_4.$$

Air is then excluded from the furnace, whereupon the sulphide reacts with the oxide and sulphate as follows:

$$2\,PbO + PbS \rightarrow 3\,Pb + SO_2,$$
$$PbSO_4 + PbS \rightarrow 2\,Pb + 2\,SO_2.$$

246. Properties and Use of Lead. Lead is the heaviest of the cheaper metals, having a specific gravity of 11.34. It has a bright white luster when freshly cut but tarnishes quickly in the air. The many uses of this metal depend chiefly upon its low melting point, its great density, and its softness. It is so soft that it can be pressed through a die into the form of tubing, which is used by plumbers to make waste pipes for sinks and sometimes even for water pipes. Sheet lead is used for roofing and for lining tanks. Lead foil forms an air-tight package for tea. Mixed with 20 per cent of antimony it furnishes type metal; with 0.5 per cent of arsenic it is used for small shot; and equal parts of lead and tin form solder. Lead dissolves easily in nitric acid, forming the nitrate $Pb(NO_3)_2$. Hydrochloric and sulphuric acids have little effect on it; hence its use in lining the chambers and pans used in manufacturing sulphuric acid.

247. Lead in Drinking Water. Lead pipes are sometimes used to convey drinking water, but there is some danger in this use of lead unless the water is hard. Hard water will generally cover lead with a coating of carbonate which is insoluble and which protects the metal from further action. Soft water, however, dissolves some of the lead, which when thus dissolved in water, acts as a poison. Cases of lead poisoning produced by water which has passed through lead pipes are not uncommon. Lead acts as a cumulative poison, so that minute quantities taken daily for some weeks or months may finally produce fatal results.

248. Oxides of Lead. Litharge, or lead monoxide (PbO), is readily formed by exposing lead at a red heat to the action of air. It is a buff-colored substance and is used in the preparation of boiled linseed oil and in the manufacture of glass and enamels. Red lead, or lead tetroxide (Pb_3O_4), is formed by heating the monoxide to about 350 degrees C. It is used in making flint glass and in the common red paint used on iron work. Mixed with linseed oil it is used in plumbing and gas fitting to make tight joints.

249. Sugar of lead ($Pb(C_2H_3O_2)_2$) is a soluble salt prepared by dissolving litharge in vinegar or acetic acid. Sugar of lead is sometimes used in hair dyes, but its use in this way is considered dangerous, as it is likely to produce paralysis.

250. White lead is a basic carbonate of lead, having the composition $2\,PbCO_3 \cdot Pb(OH)_2$. It is one of the most important compounds of lead and is used in the preparation of paints. It is prepared by the action of carbon dioxide on the oxide and acetate of lead. As a pigment for paints it has an advantage over all others in its much greater opacity and covering power. It has the disadvantage that it is readily blackened by hydrogen sulphide, and it is poisonous to the workmen who handle it. House painters are sometimes subject to a painful disease known as *lead colic,* which is caused by the slow absorption of small particles of white lead into the system. White lead is often adulterated with chalk or barium sulphate. It is being replaced for some purposes by zinc white (219), which is not discolored by hydrogen sulphide.

251. Lead arsenate ($Pb_3(AsO_4)_2$) is a white substance usually sold in the form of a paste or a powder. It is used for poisoning potato bugs and as an insecticide for use in spraying trees and shrubs. It is less soluble than Paris green and adheres more firmly to the plant, and is not so likely to injure the plant.

FIG. 125. — Testing for lead with blowpipe.

252. Tests for Lead. If a compound of lead is mixed with sodium carbonate and heated on a piece of charcoal in the inner flame of a blowpipe (Fig. 125), a small bead of metallic lead is obtained, and the softness of the bead indicates the nature of the metal. A solution of potassium chromate added to a solution containing lead gives a precipitate of chrome yellow. Hydrogen sulphide gives a black precipitate of lead sulphide (PbS).

EXERCISES

Ex. 139. What is the principal ore of lead? How is lead prepared from the ore? Examine a piece of freshly cut lead and state its properties. What happens to lead when heated? When exposed to the air? Is it easily melted and tarnished? What physical properties adapt it for its extensive use? What are some of the uses made of lead? Why is it used in lining the chambers and pans in making sulphuric acid? Is there any lead in lead pencils? Why should lead pipes be avoided for carrying drinking water? Which dissolves more lead, hard or soft water?

Ex. 140. What is the composition of the red lead used in paint? What is litharge and for what is it used? What is sugar of lead? Test a sample of hair dye from the local drug store by adding a little hydrogen sulphide. A black precipitate indicates the presence of lead.

Ex. 141. Test a lead compound according to paragraph 252. What is the result? The same test may be applied to the precipitate from the hair dye to confirm the test for lead.

Ex. 142. What is the substance known as white lead? How is it manufactured? For what is it used? What advantage and what disadvantage does it have for use in paints? Test a sample of paint for the presence of lead. Is white lead often adulterated? What advantage does zinc oxide have over white lead? Why does white lead paint become blackened?

CHAPTER XXVI

COPPER

253. Copper is, industrially, one of the most important of the metallic elements. It occurs free in nature and for this reason has, together with silver and gold, been known from a very early period. It is characterized by its reddish color. This color, however, is seen only in fresh surfaces since the metal soon becomes covered with a film of oxide, sulphide, or carbonate. Copper is flexible, hard, and tough, and can readily be drawn into wire or rolled into very thin sheets. Next to silver it is the best-known conductor of heat and electricity. Hydrochloric acid and cold sulphuric acid have little effect upon copper. It dissolves in nitric acid, forming the nitrate ($Cu(NO_3)_2$) and various oxides of nitrogen (143). With hot sulphuric acid it forms the sulphate ($CuSO_4$) and sulphur dioxide. The most common ore of copper is copper pyrites ($CuFeS_2$). Copper is precipitated from solutions of its salts by iron, zinc, and some other metals, as can be shown by dipping a piece of bright iron or steel into a solution of copper sulphate (Fig. 126). Copper will immediately be deposited as a thin, red coating upon the

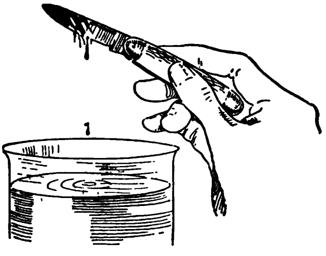

Fig. 126. — Copper deposited on a knife blade.

iron, and a corresponding amount of the iron will go into solution to replace the copper.

$$CuSO_4 + Fe \rightarrow FeSO_4 + Cu.$$

254. Uses of Copper. Next to iron, copper is the most useful metal. Enormous quantities of copper wire are used in operating the telegraph, the telephone, the electric railway, and the electric light. Sheet copper is made into household utensils, boilers, and stills. Copper bolts, nails, and rivets are used in ships because copper rust does not destroy wood as iron rust does. It is used for ornamental and artistic purposes, and in the printing trade for engraving and electrotyping.

255. Alloys. Some metals intermix when melted together and when cooled form a metal-like substance which has properties somewhat different from either of the separate metals. Such a mixture of two or more metals is called an *alloy*. Copper is a constituent part of many of the most important alloys. The following are some common alloys:

Brass	63%-73% copper,	27%-37% zinc
Bronze . . .	70%-95% copper,	1%-25% zinc, 1%-18% tin
German silver .	50%-60% copper,	20% zinc, 20%-30% nickel
Gun-metal . .	90% copper,	10% tin
Silver coin . .	10% copper,	90% silver
Nickel coin . .	75% copper,	25% nickel
Bell metal . .	75% copper,	25% zinc.

256. Copper-plating and Electrotyping. Copper is deposited from a solution of its salts by the electric current.

This fact is utilized in copper-plating. The object to be plated is connected with the negative pole of a battery and hung in a solution of copper sulphate. The other pole is connected with a bar of copper which is also immersed in the solution. As the current passes through the solution, the copper salt is decomposed, and the copper is deposited on the object to be plated. A like amount of copper is dissolved from the bar connected with the positive pole.

Books are often printed from electrotype plates. These are made by first taking an impression of the type in wax. The inside of the mold thus formed is dusted over with powdered graphite in order to make it conduct electricity. The mold is connected with the negative pole of the battery and suspended in a solution of copper sulphate. As the current passes, copper is deposited upon the mold in a coherent film, and a perfect copy of the type is obtained. The sheet is strengthened by filling in the under surface with melted lead. For the daily newspaper this process is too slow and the printing is done from stereotype plates, which are made by pouring melted stereotype metal, consisting of lead, antimony, and tin, into a paper pulp cast of the type. This process gives a printing surface much inferior to the electrotype plate.

257. Oxides of Copper. When copper is heated to redness in the presence of plenty of air or oxygen, a black oxide is formed having the formula, CuO. This compound is called cupric oxide. In the absence of sufficient oxygen to form the cupric oxide, another oxide having a bright red color is produced. This is cuprous oxide, Cu_2O. Two series of salts (cupric and cuprous) of copper can be prepared corresponding to these two oxides, but only the cupric compounds are of any great commercial importance.

258. Copper sulphate, $CuSO_4$, is the best known of the salts of copper. It crystallizes from water in beautiful blue crystals having the formula $CuSO_4 \cdot 5\ H_2O$, commonly called *blue vitriol* or *bluestone.* When heated, the crystals lose water and the mass becomes white. The colorless substance becomes blue again in contact with water. It is used in galvanic batteries, in copper-plating, and as a starting point for the formation of various compounds of copper. All copper compounds are poisonous to plants and animals.

259. Copper hydroxide, $Cu(OH)_2$, is formed as an insoluble precipitate when any soluble hydroxide is added to a solution of copper sulphate or to any other soluble salt of copper:

$$CuSO_4 + Ca(OH)_2 \rightarrow Cu(OH)_2 + CaSO_4.$$

It is light blue in color. Copper hydroxide formed according to the above reaction from copper sulphate and slaked lime is known as *Bordeaux mixture* and is used in spraying plants for the control of certain fungous diseases. Bordeaux mixture alone is a fungicide merely and has little poisonous effect on insects. To make it an insecticide as well arsenate of lead (251) is commonly added to the mixture. Ammonia water when added to a solution of copper sulphate first precipitates copper hydroxide and then redissolves it, forming a deep blue solution. This preparation has been used as a fungicide under the name of *ammoniacal copper sulphate,* but its use is largely superseded by Bordeaux mixture.

260. Paris green is a complex compound of copper with arsenious acid and acetic acid. It is a brilliant green material sometimes called emerald green. It is used in paint making, but more largely for the destruction of potato bugs and other injurious insects. It is slightly soluble and if used in too

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negative. The negative is placed on the paper and exposed to the light in such a way that light will pass through the negative, which obstructs the light in proportion to the thickness of the silver deposit. The print, therefore, is the reverse of the negative and has the same shading as the object photographed. The print may be developed, fixed, and washed in the same manner as the negative.

267. Blue prints are produced by the action of light on a salt of iron. A solution of ferric ammonium citrate and potassium ferricyanide is brushed on a sheet of paper and dried in the dark. If a design drawn on tracing cloth, or a photographic negative, is placed over this paper and the two are exposed to the sunlight, the sensitive paper turns to a brownish color where the light penetrates the negative. Under the black parts of the negative where no light strikes it the paper is unaffected. If, now, the paper is washed in water the unchanged iron salt is removed, while that part affected by the light turns to a blue color. The blue color is due to the formation of the insoluble salt of iron known as Prussian blue. Blue print paper is used in large quantities by architects, engineers, and designers.

268. The test for silver is the presence of the white curdy precipitate of silver chloride formed by the addition of hydrochloric acid to a silver salt. This precipitate is soluble in ammonia water and darkens when exposed to light.

EXERCISES

Ex. **148.** Give the physical properties of silver. Is it readily oxidized? Why does it turn black when exposed to the air? When in contact with egg, or rubber, or perspiration? Dissolve a ten-cent piece in nitric acid. What does it contain besides silver? What is the percentage composition of American silver coins? Of British coins? What is meant by sterling silver? How are other metals silver-coated?

Ex. 149. What is the name and formula of the salt produced by the action of nitric acid on silver? What is the common name for it? For what is it used? Take a little of the solution of the coin and add ammonia water until the precipitate which first forms is again dissolved. With a clean pen write your name on a piece of white cloth with this liquid. When dry press a hot iron on the writing. What happens? This illustrates one method of making indelible ink.

Ex. 150. Dilute some of the above solution of the coin with water in a test tube and add a little common salt. What happens? Write the reaction. Expose the test tube and contents to strong sunlight and note what happens. Paint a piece of plain white paper with a weak solution of salt and allow it to dry. Now brush it over with a solution of silver nitrate (in a dark room). Place a fern leaf or other object on the paper and expose it to the sunlight, using a pane of glass to keep the leaf pressed against the paper. What happens? How is this property of silver salts utilized in photography?

Ex. 151. (Teacher) Expose and develop a photographic plate so that the class may watch the different steps. If a very slow plate such as a lantern slide plate is used, the development may be done in any well-darkened room, in case a regular dark room is not available. When the negative is dry make a print with any good developing paper. Use developers recommended by the manufacturers. What does the light coming through the lens do to the plate? What does the developer do? Why is the hypo used? What would happen if the plate were not fixed? Explain the different steps in the production of the print.

Ex. 152. Dissolve 1 gram potassium ferricyanide in 5 cc. water; in a separate test tube, dissolve 1½ grams ferric ammonium citrate in 5 cc. water and mix the two solutions. Now paint a piece of paper with the solution in a dark room by candlelight and let the paper dry. Expose the paper to the sun's rays under a design on tracing cloth, or a photographic negative. When the exposure has continued long enough (five or ten minutes, according to the amount of light) wash the paper. What change has taken place? What effect did the light have on the iron salt? Tell about any blue prints which you have seen.

Ex. 153. To a little coin solution add hydrochloric acid. What occurs? Add an excess of strong ammonia water. Results? This illustrates the test for silver, as no other metal forms a chloride that is insoluble in nitric acid and water and is soluble in ammonia water.

CHAPTER XXVIII

REVIEW OF THE METALLIC SALTS — RECOGNITION OF THE COMMON METALS

269. IN the foregoing chapters ten of the more important metals have been briefly discussed. Of these not more than seven are commonly met with in the metallic form; the others are rarely seen outside of the chemical laboratory. Only three or four of the metals are found in nature in the free state; the others occur in the oxides, sulphides, carbonates, or silicates. A large number of compounds of the metals are of commercial importance, notably the salts, and a few of the oxides and hydroxides.

270. Metallic Salts. Theoretically at least, every metal should be able to combine with every acid to form a salt, and as there are many acids and metals the possible number of salts is very large. Only the more important of these salts and especially those which are of more or less common use have been mentioned in this text. Descriptions of many others will be found in the larger texts on chemistry.

271. Preparation of Salts. There are several general methods which may be used for the preparation of salts.

(1) A salt may be formed by the action of the proper acid on a metal, an oxide, hydroxide, or carbonate:

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + 2\,H,$$
$$CaO + 2\,HCl \rightarrow CaCl_2 + H_2O,$$
$$KOH + HNO_3 \rightarrow KNO_3 + H_2O,$$
$$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2.$$

(2) If a salt is insoluble in water it may be prepared by precipitation. To a solution of a soluble salt of the metal is added a solution of a soluble salt of the desired acid. For example, calcium carbonate may be prepared by adding a solution of sodium carbonate to a solution of calcium chloride, whereupon the calcium carbonate will be precipitated:

$$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2\,NaCl.$$

If the desired salt is insoluble in acids, it may be precipitated by adding the desired acid itself in place of the salt:

$$AgNO_3 + HCl \rightarrow AgCl + HNO_3.$$

(3) Some of the binary salts may be formed by the direct union of the metallic and non-metallic elements:

$$Fe + 3\,Cl \rightarrow FeCl_3,$$
$$Cu + S \rightarrow CuS.$$

272. Solubility of Salts. A knowledge of the solubility in water of the different salts is of importance when devising a method for the preparation of a salt. It is an aid also in the determination of the basic and acid parts of the salt. The solubility of the salts and of the hydroxides of the metals studied may be summarized as follows:

(1) *Hydroxides* are insoluble except those of ammonium, sodium, potassium, and calcium.

(2) *Nitrates* are all soluble.

(3) *Chlorides* are soluble except silver chloride. Lead chloride is slightly soluble in cold water and quite soluble in hot water.

(4) *Sulphates* are soluble except those of lead and calcium. Calcium sulphate is slightly soluble.

(5) *Sulphides* are insoluble except those of ammonium,

sodium, potassium, calcium, and magnesium. The insoluble sulphides can be divided into two classes: (*a*) Those soluble in dilute acid, as lead and copper. (*b*) Those insoluble in dilute acid, as iron, aluminum, and zinc.

(6) Carbonates, phosphates, and silicates are insoluble except those of ammonium, sodium, and potassium.

273. Recognition of the Metals. In Chapter XX was given a scheme for the recognition of the acid radicals of salts. Consideration is now given to the methods of recognizing the basic or metallic part of a salt. The methods given here apply only to single substances, as the analysis of a mixture of substances calls for a knowledge of analytical chemistry. The tests will be confined to the ten of the more common metallic elements studied: sodium, potassium, calcium, magnesium, zinc, aluminum, iron, lead, copper, and silver.

(1) If the substance is one of those studied in connection with Chapter XX and the acid radical has been determined, some idea as to the metal may be gained by comparing with the solubilities given above. For example, if the acid is found to be carbonic or phosphoric acid and the substance dissolves in water, the base must be ammonium, sodium, or potassium; if the substance is a chloride and is insoluble in water, it must be silver chloride. An insoluble sulphate must be lead or calcium, and so on.

(2) The usual method of detecting the metals is by the use of reagents that precipitate certain groups of metals, which may again be subdivided until the metal is traced. The substance to be tested is dissolved in water. If it is not soluble in pure water the least amount of nitric acid should be used, and the following procedure carried out.

(*a*) Add to the solution a little hydrochloric acid. If a precipitate is formed, it is silver chloride (268), or lead chloride. Add ammonia water, which will dissolve silver chloride. Lead chloride if present is not dissolved by ammonia water but will dissolve if the liquid is heated. Confirm by test (252).

(*b*) If no precipitate is produced by hydrochloric acid, take half of the solution and pass hydrogen sulphide gas through it. A precipitate may be copper sulphide or lead sulphide. The first solution may have been too dilute for the lead to precipitate as chloride. Determine which is present by testing original substance for lead (252) and copper (261).

(*c*) If no precipitation is caused by hydrogen sulphide, make the solution alkaline by adding ammonia water. A brownish precipitate is iron (244), while a white precipitate indicates aluminum (230) or zinc (221).

(*d*) If none of these reagents gives a precipitate, evaporate the reserved half of solution to a small bulk and add a few drops of sulphuric acid. A precipitate indicates calcium.

(*e*) If no precipitate is produced by the sulphuric acid, make the solution alkaline with ammonia water and add a solution of sodium phosphate. A precipitate indicates magnesium. Confirm by testing the original substance (216).

(*f*) If none of these reagents produces a precipitate, the substance is a salt of sodium or potassium. Determine which by the flame test (131 and 210).

EXERCISE

Ex. 154. Obtain from the teacher samples of simple chemical compounds to test for the metal or basic radical. Read this chapter carefully and follow the outline exactly. The metal will be one of the following: Sodium, Potassium, Calcium, Magnesium, Copper, Silver, Lead, Zinc, Aluminum, or Iron. Make a record of the result of each test.

To the Teacher. This chapter should be made the basis of a thorough review of the chemistry of the metals.

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The first four members are gases at ordinary temperatures. From pentane to pentadecane ($C_{15}H_{32}$) they are liquids with higher boiling points, and from hexadecane upward they are solids with increasingly higher melting points.

277. Petroleum is the principal source of the hydrocarbons of the methane series. It is a thick, greenish-brown oil found in oil-bearing strata of the earth. The chief oil-producing areas of the United States are in Oklahoma, California, Texas, Illinois, Louisiana, West Virginia, Pennsylvania, and Ohio. Petroleum is pumped from wells sunk in the ground, and is stored in large tanks or conveyed directly to the refineries. Sometimes when the oil is under great pressure of gas, the newly driven well spouts oil from the surface. Such a well is called a gusher (Fig. 130). Petroleum is a very complex mixture of hydrocarbons, and while some of it is used in the crude condition as a fuel, its principal value lies in the fact that it is the source of many useful products such as gasoline, vaseline, and paraffin. The crude petroleum is placed in large stills and is subjected to distillation. The portion of the liquid that distills between the temperatures of 70 degrees F.

FIG. 130. — A spouting oil well or gusher.

and 150 degrees is called naphtha; that between 150° and 300° is kerosene; while that which distills between 300° and 400° is used for lubricating oil. When the remaining oil is chilled the solid constituents separate and constitute ordinary paraffin. Naphtha is again separated into petroleum ether, gasoline, and benzine. In some refineries a semisolid fraction is also obtained which is the vaseline, or petrolatum, of commerce.

None of these substances is a single chemical compound, but each one is composed of several hydrocarbons. Gasoline consists of hydrocarbons with low boiling points, while paraffin contains those with very high boiling points.

278. Gasoline. The chief uses of gasoline depend upon the fact that it is very volatile and is, therefore, easily converted into a gas. The mixture of gasoline vapor and air is explosive, and this quality is utilized in the gasoline engine. When gasoline is used in heating or illuminating, it is first converted into a gas. This is accomplished by heating it or by forcing air through it. The volatile character of gasoline is also the cause of many accidents. It is to be noted that nearly all the accidents with gasoline stoves have been caused by the fact that the gasoline tanks were filled while the burner was lighted. Gasoline is one of the best solvents for fats and its use in cleaning depends upon this fact.

279. Acetylene (C_2H_2) is a colorless gas now extensively used for illumination. It is prepared by the action of water on calcium carbide (169):

$$CaC_2 + 2\,H_2O \rightarrow Ca(OH)_2 + C_2H_2.$$

A special form of burner (Fig. 131) is required in burning acetylene to prevent the formation of soot. When the gas is used in such a burner the flame is very white and brilliant. Acetylene is readily decomposed, and it was formerly dangerous to handle when compressed in cylinders. It has been discovered that if the cylinder is filled with a porous material like asbestos and this material is saturated with a compound called acetone, acetylene may be forced into the cylinder under high pressure with perfect safety. Acetylene, when burned with oxygen in an apparatus much like the oxyhydrogen blowpipe, makes the hottest known flame (2700°). It is sometimes used in cutting iron, as it melts its way through the iron at the point of contact.

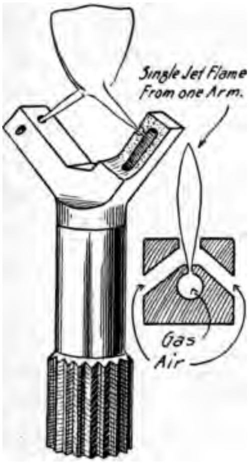

FIG. 131.—**Acetylene burner.**

EXERCISES

Ex. **155.** Mix thoroughly a teaspoonful each of sodium acetate and soda-lime. Place in a test tube, arrange as in Fig. 50, and heat. Collect two bottles full of the gas. What is the composition of this gas? The name? The reaction? Ascertain whether the gas will burn. Fill a bottle half full of the gas and half full of air and ignite it. Does this mixture explode? Where is this gas found in nature? What do miners call it? If there is a marsh or a pond near the school where leaves and other organic matter are decomposing, try to collect some methane as shown in Fig. 129. Why is this gas commonly called marsh gas?

Ex. **156.** What is meant by a hydrocarbon? Are many combinations of carbon and hydrogen known? As the number of carbon atoms in the molecule increases what effect does it have on the boiling point of the hydrocarbon? Are the hydrocarbons with the largest molecules liquids or solids?

Ex. 157. Examine a sample of crude petroleum; also samples of gasoline, kerosene, and lubricating oil. What is the original source of all these substances? What is the chemical nature of petroleum? Are gasoline and kerosene single chemical compounds? In which do the hydrocarbons have the larger molecules? What are some of the important uses of gasoline? For what is it used at your home?

Ex. 158. Place some calcium carbide in an apparatus as shown in Fig. 132. The bulb (*B*) contains water and is fitted with a stopcock so that the water can be slowly run into the flask (*A*). When ready to collect the gas allow the water to run into the flask a drop at a time. Examine the gas which is evolved. What is its name and formula? Give its reaction of formation. Why must it be burned in a special burner? Are any houses in your neighborhood lighted with acetylene? What gas is in the gas tanks used on automobiles and motorcycles? What temperature is obtained where acetylene is burned with oxygen? What practical application is made of this intense heat?

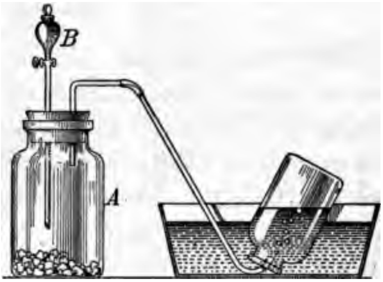

FIG. 132. — **Laboratory apparatus for the production of acetylene.**

CHAPTER XXX

ALCOHOLS

280. Wood Alcohol. When wood is heated in closed retorts in the manufacture of charcoal (90), a distillate consisting of a number of substances is obtained. One of these compounds is *wood alcohol* (methyl alcohol), which has the formula CH_3OH. Wood alcohol is a colorless liquid that boils at 65° C. and burns with a colorless and sootless flame. It is a good solvent for resins and is used in making certain varnishes and shellacs. It is burned also in alcohol lamps, although its use for this purpose and for varnishes has decreased since the introduction of denatured alcohol. Wood alcohol is poisonous and it produces paralysis of the optic nerve. Many cases of blindness have been caused by the drinking of cheap whiskies adulterated with wood alcohol, and by the continued inhaling of its vapor.

281. Ordinary alcohol, also known as grain alcohol and *ethyl alcohol,* has the formula C_2H_5OH. It is obtained from the fermentation of sugars. The fermentation is brought about by the action of yeast. In the case of grape sugar ($C_6H_{12}O_6$) the reaction may be represented as follows:

$$C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2.$$

Alcohol is prepared commercially from substances rich in starch, such as corn or potatoes. The starch is first converted into a sugar by means of malt, and yeast is then added.

Yeast is a microscopic vegetable organism, which during its growth produces a number of changes resulting in converting the sugar into alcohol. The resulting alcohol is separated from the fermented liquid by distillation (Fig. 133). The alcohol of commerce contains about 5 per cent of water.

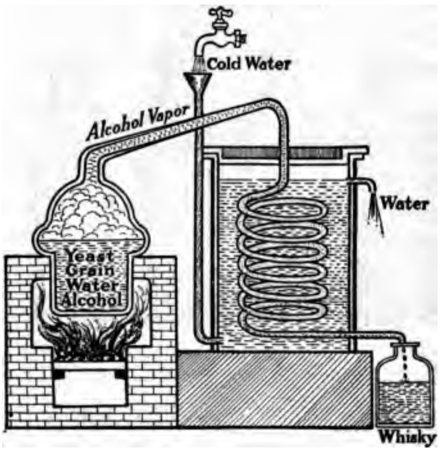

FIG. 133. — A still for the production of alcoholic liquors.

Alcohol is a colorless liquid with a characteristic, pleasant odor. It boils at 78° C. It is sometimes used as a fuel, especially in spirit lamps, as it burns with a colorless and sootless flame. It burns according to the following reaction:

$$C_2H_5OH + 6\,O \rightarrow 2\,CO_2 + 3\,H_2O.$$

It is a solvent for many substances. Pharmacists use it in the preparation of tinctures, essences, and extracts. Many of the better grades of varnishes and shellacs contain ethyl alcohol. When the term alcohol is used without a qualifying word, ethyl alcohol is always meant. When taken into the system in small quantities, alcohol produces intoxication; in larger amounts it acts as a more positive poison.

282. Alcoholic Beverages. Many beverages contain alcohol in greater or smaller quantity (Fig. 134). In all cases the alcohol is produced by fermentation. Wines are made by the fermentation of the sugars in fruit juices, particularly of the grape. Wines contain 5 to 15 per cent of alcohol. Hard cider is really a wine produced from apple juice. The yeast plant, since it is always associated with the fruits, need not be added in wine making. Beer is made by the fermentation of malt; in this case the yeast is added. Beer contains from 3 to 5 per cent of alcohol. Corn, rice, and glucose are sometimes used to replace part of the malt.

FIG. 134. — The percentage of alcohol in distilled liquor, wine, and beer.

Whisky, brandy, rum, and gin are known as *distilled liquors*. They contain from 40 to 60 per cent of alcohol. Whisky is made by distilling a beer made from rye, corn, or barley. Brandy is made by distilling wine, or the fermented juice of apples, peaches, cherries, or other fruits. Rum is distilled from the liquid obtained by fermenting molasses. Gin is an alcoholic liquor flavored with oil of juniper berries.

283. Denatured alcohol is ethyl alcohol, to which is added wood alcohol, benzine, or a bad-smelling compound prepared by heating bones and known as pyridine. These substances make its use for beverages or medicine impossible. About four fifths of the cost of ordinary alcohol is due to the internal revenue tax imposed by the government. Denatured alcohol is tax free, to encourage its use in the arts. The denaturing in no way impairs its value as a fuel or for use in varnishes and shellacs.

284. Alcohols Are Bases. The formulas for methyl alcohol (CH_3OH) and ethyl alcohol (C_2H_5OH) are written in such a way as to show the presence of hydroxyl in the compound. This is done because the alcohols are all bases in the same way that ammonium hydroxide (NH_4OH) is a base, although the basic character of the alcohols is less pronounced. They react with acids in a manner similar to ammonium hydroxide:

$$NH_4OH + HCl \rightarrow H_2O + NH_4Cl \text{ (ammonium chloride)};$$
$$C_2H_5OH + HCl \rightarrow H_2O + C_2H_5Cl \text{ (ethyl chloride)}.$$

Methyl and ethyl alcohols are hydroxides made by replacing one hydrogen atom of methane and ethane with hydroxyl. Similar alcohols can be prepared corresponding to the more complex hydrocarbons, but most of them are of minor importance.

285. Glycerin. Just as there are inorganic bases with more than one hydroxyl group in the molecule, as, for example, $Ca(OH)_2$ and $Al(OH)_3$, so there are alcohols with more than one hydroxyl group. The most important of these is glycerol, known commercially as *glycerin*, $C_3H_5(OH)_3$. Glycerin is a heavy, colorless, sirupy liquid with a sweet taste. It is miscible with water and is so hygroscopic that it will absorb half its weight of water from the moisture of the air. It is used in cosmetic and medicinal preparations, in ink rollers of printers, in the ink for rubber stamps, and to soften leather. Glycerin is one of the products obtained during the manufacture of soap (303), and in the preparation of the stearin used in making candles.

286. Nitroglycerin is trinitrate of glycerin, made by slowly adding glycerin to a mixture of nitric and sulphuric acids:

$$C_3H_5(OH)_3 + 3\ HNO_3 \rightarrow C_3H_5(NO_3)_3 + 3\ H_2O.$$

The sulphuric acid does not take part in the reaction, but causes the action to continue by keeping the mixture dehydrated. Nitroglycerin is a heavy, colorless, oily liquid. It explodes when heated to 180° C., or when subjected to a shock. Because of the danger in handling pure nitroglycerin, it is mixed with some inert, porous substance, such as infusorial earth or wood pulp. This mixture is called *dynamite,* and the different grades are classified and named according to the percentage of nitroglycerin they contain.

287. Formaldehyde. When methyl alcohol is burned in a limited supply of air, or the mixture of air and the vapor of methyl alcohol is passed over heated copper, a gas is formed which is known as formaldehyde (CH_2O):

$$CH_3OH + O \rightarrow CH_2O + H_2O.$$

Formaldehyde is a gas with a stinging, stifling odor which causes the eyes to smart. It is a powerful germicide and is largely used to disinfect buildings following cases of contagious diseases. It is more effective than sulphur dioxide and has no bleaching effect (63). An aqueous solution containing 40 per cent of the gas is sold under the name of formalin. This is used also as a disinfectant, and for the treatment of seed potatoes for the destruction of scab, and of oats and other grain to destroy smut. Formaldehyde is used also for the preservation of anatomical specimens and to harden gelatin films in photography. It is sometimes improperly employed as a food preservative.

EXERCISES

Ex. **159.** (Teacher) Place a cupful of commercial glucose, common molasses, or Karo sirup, in a two quart bottle (*A*) and add a quart of lukewarm water. Rub a cake of compressed yeast in half a cupful of water and add it to the mixture in the bottle. Connect the large bottle with a small bottle (*B*) containing limewater, as shown in Fig. 135. Set this aside in a warm place for two or three days. Bubbles of gas will soon form in *A* and pass into *B*. What happens to the limewater? What is the gas formed in *A*? After fermentation has ceased, decant about half of the liquid in *A* into a distilling flask.[1] Connect with a condenser and distill off 10 to 15 cc. Compare the odor with the alcohol of the laboratory. Put 2 or 3 cc. in an evaporating dish and test it with the flame. Will it burn?

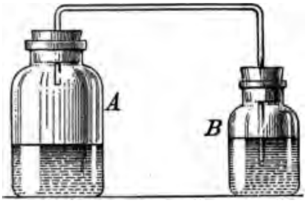

FIG. 135. — **Producing alcohol in the laboratory.**

Ex. **160**. State the properties of ordinary alcohol. What is its chemical name? How is it made? Write the reaction for change of glucose to alcohol. Name some beverages that contain alcohol. What uses are made of alcohol? What is meant by denatured alcohol? Why is denatured alcohol so much cheaper than the pure alcohol? Does denaturing impair its value for use in the arts?

Ex. **161**. How is wood alcohol prepared? What is the chemical name for wood alcohol? For what is it used? Why is caution necessary in using it? Why are alcohols known as bases?

Ex. **162**. Examine a sample of glycerin and give its properties; its formula. For what is it used? What is the source of commercial glycerin? How is nitroglycerin made? Write the reaction. What is dynamite?

Ex. **163**. Examine a sample of formaldehyde. How is it prepared? Give its formula. For what is it used? Why is it preferable to sulphur dioxide for disinfecting after sickness?

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[1] Set the bottle *A* aside and keep it open for two or three weeks to see whether acetic acid will develop.

CHAPTER XXXI

ORGANIC ACIDS

288. Acetic Acid. It will be recalled that apple juice, upon standing, undergoes a fermentation that results in the formation of alcohol. Hard cider may contain from 4 to 8 per cent of alcohol. Upon longer standing the cider becomes very sour and is then called *vinegar.* The sour taste of the vinegar is due to the fact that the alcohol has been changed into a substance known as *acetic acid* ($H \cdot C_2H_3O_2$):

$$C_2H_5OH + 2\,O \rightarrow H \cdot C_2H_3O_2 + H_2O.$$

This acid is the most familiar member of a class of compounds known as *organic acids.* All of them resemble the inorganic acids in their general behavior, but are weaker acids. *Only one of the hydrogen atoms of acetic acid has acid properties,* that is, can be replaced by a metal; and that fact is indicated in the formula by separating the replaceable hydrogen from the rest of the molecule. This method of indicating the replaceable hydrogen atoms of the organic acids is used throughout this text.

The change of alcohol into acetic acid during the formation of vinegar is brought about by the action of a species of bacteria known as *Bacterium aceti,* and the change is known as *acetic fermentation,* which is evidently an oxidation process. The slimy substance sometimes found in vinegar and called *mother of vinegar* consists of masses of these bacteria. Vinegar contains from 4 to 6 per cent of acetic acid.

The old method of producing vinegar from cider, in which the fermentation was allowed to take place in barrels, required many weeks or months, as the oxidation could take place only at the surface of the liquid. In the method known as the *quick vinegar process* (Fig. 136), tall barrels are loosely filled with beech-wood shavings, which are then moistened with old vinegar to introduce the bacteria. The cider or other liquid containing alcohol is allowed to trickle slowly over the shavings and is thus exposed to the action of bacteria and to the oxygen of the air. By this process the vinegar is produced in about ten days. Vinegars are sometimes made from wine, and from fermented malt extract (malt vinegar). *White wine vinegar*, or *distilled vinegar*, is made by treating solutions of pure alcohol by the process just described.

FIG. 136. — **Vinegar-making by quick process.**

289. Acetic Acid from Wood. Acetic acid is one of the products of the destructive distillation of wood (90). It was formerly called " pyroligneous acid "; that is, the acid made by heating wood. Much of the acetic acid on the market is made in this way, and some of the cheap vinegar is merely a 4 per cent solution of this acid.

Acetic acid, when pure, is a colorless liquid with a pungent odor. It solidifies at 17° C. and is soluble in all proportions in water. With metals it forms salts, among the most important of which are sodium acetate ($NaC_2H_3O_2$), and lead acetate ($Pb(C_2H_3O_2)_2$) commonly called *sugar of lead.* Copper acetate ($Cu(C_2H_3O)_2$) when combined with copper arsenite forms Paris green (260).

290. Lactic Acid. Milk upon standing gradually becomes sour because of the formation of lactic acid ($H \cdot C_3H_5O_3$). This acid is produced by the action of lactic-acid bacteria upon milk sugar (310). This process is known as *lactic fermentation.* Lactic acid may also be prepared by the fermentation of other sugars, and as it is now of some industrial importance it is produced by the action of the bacteria found in old cheese, upon solutions of glucose, or cane sugar. When pure, it is a colorless, sirupy liquid with an intensely sour taste.

291. Oxalic acid ($H_2 \cdot C_2O_4$) exists in many plants. It gives the sour taste to sour grass and to sheep sorrel. It differs from most of the acids so far studied in being a solid instead of a liquid. It forms large, colorless crystals containing two molecules of water of crystallization ($H_2 \cdot C_2O_4 \cdot 2\,H_2O$). Oxalic acid is poisonous. Its antidote is calcium carbonate (chalk), which forms with it the insoluble calcium oxalate (CaC_2O_4). Calcium oxalate is found in the clovers and many other plants. Oxalic acid is useful in removing ink and rust spots from floors and fabrics. It is used also in cleaning brass, and in bleaching straw hats.

292. Tartaric acid ($H_2 \cdot C_4H_4O_6$) is the acid of grapes. When pure it forms beautiful, large, prismatic crystals which are readily soluble in water. In grapes it is found as acid potassium tartrate ($KH \cdot C_4H_4O_6$). This is the substance commonly called *cream of tartar.* When a solution of cream of tartar is neutralized with sodium hydroxide, Rochelle salt ($KNaC_4H_4O_6$) is formed:

$$KHC_4H_4O_6 + NaOH \rightarrow KNaC_4H_4O_6 + H_2O.$$

293. Citric acid ($H_3 \cdot C_6H_5O_7$) occurs in lemons, oranges, and other citrus fruits. It exists also in currants, gooseberries, and cranberries. Citric acid forms large colorless crystals which are soluble in water. Both citric and tartaric acid are often used as substitutes for lemons in making cheap lemonade and other acidulated beverages. Magnesium citrate ($Mg_3(C_6H_5O_7)_2$) is employed as a purgative in medicine.

294. Tannic acids, or **tannins,** are substances with an astringent taste. They are widely distributed in the vegetable kingdom. The principal commercial sources are the bark of the oak and hemlock, sumach, nutgalls, and a number of Indian and South American trees. Tannins occur also in smaller quantities in the leaves of many plants. Tannins are used for tanning leather, in dyeing, and in making ink.

Ink is the black mixture obtained by mixing solutions of tannin with iron salts. One recipe for black ink is as follows: Extract 100 grams of powdered nutgalls with 1.4 liters of water, and add 50 grams of gum arabic and 50 grams of ferrous sulphate. When the mixture is exposed to the air, a permanent black color is developed. A black or a blue dye is commonly added to give the ink a temporary color.

The operation of tanning, or the conversion of animal skin into leather, depends on the formation in the skin of an insoluble compound of tannin and the albuminoid matter of the skin. The tannin is derived from oak or hemlock bark, which is ground to a coarse powder and piled in layers with the skins in deep vats. The vats are filled with water, and the skins are allowed to soak for a few weeks or months.

295. Benzoic acid ($H \cdot C_7H_5O_2$) **and salicylic acid** ($H \cdot C_7H_5O_3$) are both solid, crystalline acids that are used in medicine and sometimes as preservatives in food. Benzoic acid exists naturally in gum benzoin, while a compound of salicylic acid is the principal ingredient of oil of wintergreen. Both these acids are now prepared artificially from coal tar compounds.

Sodium benzoate ($NaC_7H_5O_2$), is frequently used as a preservative in foods. Its use is unnecessary, however, and should be discouraged. Sodium salicylate also is sometimes used as a preservative. Both it and salicylic acid are used in medicine, notably in the treatment of rheumatism.

296. Organic Salts, or Esters. If to a mixture of ethyl alcohol and acetic acid in a test tube a little sulphuric acid is added and the whole is gently heated, a vapor with a pleasant, fragrant odor is given off. This body, ethyl acetate ($C_2H_5 \cdot C_2H_3O_2$), is formed by replacing the acid hydrogen of acetic acid by the basic ethyl radical. The equation is

$$C_2H_5OH + H \cdot C_2H_3O_2 \rightarrow C_2H_5 \cdot C_2H_3O_2 + H_2O.$$

The sulphuric acid absorbs the water as fast as it is formed and permits the reaction to continue. Such salts as ethyl acetate, in which both the basic and the acid parts are organic radicals, are called *ethereal salts* or *esters,* the shorter name being preferred. Many of these esters exist naturally in fruits and impart to them their characteristic flavors. They can be prepared artificially by a process analogous to that discussed for ethyl acetate and many of the flavoring extracts on the market consist of such artificially prepared esters. Oil of wintergreen is methyl salicylate ($CH_3 \cdot C_7H_5O_3$). Most of the oil of wintergreen on the market is artificial.

When an ester is gently heated with an alkali, the alkali salt of the acid is formed and the alcohol is set free:

$$\underbrace{C_2H_5 \cdot C_2H_3O_2}_{\text{ethyl acetate}} + NaOH \rightarrow \underbrace{NaC_2H_3O_2}_{\text{sodium acetate}} + \underbrace{C_2H_5OH.}_{\text{ethyl alcohol}}$$

The esters are also decomposed by heating with steam under pressure into the acid and alcohol:

$$\underset{\text{ethyl acetate}}{C_2H_5\cdot C_2H_3O_2} + H_2O \rightarrow \underset{\text{alcohol}}{C_2H_5OH} + \underset{\text{acetic acid}}{H\cdot C_2H_3O_2}.$$

EXERCISES

Ex. **164.** What gives the sour taste to vinegar? How is vinegar made commercially? What is meant by the quick vinegar process? Write the reaction for change of alcohol to acetic acid. Give the properties of acetic acid; the formula. How is acetic acid prepared from wood? What was the old name for it? Make a four per cent solution of acetic acid and compare the flavor with cider vinegar. Is the flavor of a good vinegar entirely due to the acetic acid?

Ex. **165.** What acid is formed when milk sours? Can you prove that an acid is present? Examine the laboratory sample of lactic acid. Add a drop to a tablespoonful of water and taste it.

Ex. **166.** Examine some crystals of oxalic acid. Where is it found in nature? Try cleaning an old straw hat with a solution of oxalic acid. Rub the solution on with a sponge or a piece of cloth and place the hat in the sun.

Ex. **167.** Examine some cream of tartar. What is the source of cream of tartar? The formula? Examine crystals of tartaric acid. Give its formula. What is formed when cream of tartar is neutralized with sodium hydroxide?

Ex. **168.** To what acid is the sour taste of lemons due? What is the appearance of citric acid? In what other fruits is it found? How is cheap lemonade made?

Ex. **169.** Examine some tannic acid. Dissolve a little in water and taste a drop. Does it have an astringent taste? Where is it found in nature? Add a little tannic acid solution to a solution of ferric chloride. What is the result? Steep a little oak bark in water and add ferric chloride to the solution. Do the same with a little tea. Have you any evidence that oak bark and tea contain tannic acid? How is ink made? Why does it get darker when exposed to the air? How is tannic acid used in tanning leather? Soak a piece of lean meat in a strong solution of tannic acid for a few days. What is the condition of the meat?

Ex. 170. Examine bottles of factory-made catsup or other foods at home or in the local store and see whether any have stated on the label that sodium benzoate was used in their manufacture. Do you think any preservative should be used in canning vegetables or fruits?

Ex. 171. Place 1 cc. each of alcohol and acetic acid in a test tube. Add 2 cc. of sulphuric acid and warm the mixture. Note the pleasant odor evolved. What is this substance? What is the general name for such compounds? How are they sometimes used?

Ex. 172. Place a little wood alcohol and some salicylic acid in a test tube. Add sulphuric acid and warm. What odor is given off? What is the chemical composition of oil of wintergreen? What happens to esters when heated with an alkali? When heated with steam under pressure?

CHAPTER XXXII

FATS, OILS, AND SOAPS

297. Fats and oils are the products of both vegetable and animal life. Oils are liquid fats. The more common animal fats are tallow, lard, and butter. Olive oil, palm oil, cottonseed oil, and linseed oil are good examples of vegetable fats. These fats are all insoluble in water. They are soluble in ether, chloroform, carbon bisulphide, gasoline, and benzine. Hence water will not remove grease spots from clothing, but benzine and other solvents of fats will.

298. Composition of Fats. When a fat is heated with steam under pressure so as to get a temperature of about 200° C. it is decomposed with the formation of glycerin (285) and one or more organic acids. The acids more commonly found in fats are palmitic acid ($H \cdot C_{16}H_{31}O_2$); stearic acid ($H \cdot C_{18}H_{35}O_2$); and oleic acid ($H \cdot C_{18}H_{33}O_2$). These and other acids found in fats are collectively known as *fatty acids*. Fats, then, are evidently esters in which the alcohol is glycerin and the acid is one of the so-called fatty acids. As glycerin has three hydroxyl groups, its ester with stearic acid must have the formula $C_3H_5(C_{18}H_{35}O_2)_3$. (Compare with glycerin nitrate (286).) The hydrolysis of glycerin stearate may be represented as follows:

$$\underset{\text{glycerin stearate}}{C_3H_5(C_{18}H_{35}O_2)_3} + 3\ H_2O \rightarrow \underset{\text{glycerin}}{C_3H_5(OH)_3} + \underset{\text{stearic acid}}{3\ H \cdot C_{18}H_{35}O_2}.$$

The glycerin of commerce and many of the fatty acids used in candle and soap making are produced in this way.

The names of the different fats are derived from those of the fatty acids found in them, by changing the *ic* of the acid to *in*. Thus the glycerin ester (or fat) of palmitic acid is called palmatin ($C_3H_5(C_{16}H_{31}O_2)_3$); of stearic acid is stearin (C_3H_5 $(C_{18}H_{35}O_2)_3$); of oleic acid, olein ($C_3H_5 \cdot (C_{18}H_{33}O_2)_3$). Stearin and palmitin are solids, while olein is a liquid. The natural fats are nearly always mixtures of the different single fats, and the consistency of the fat depends upon the proportion of the different esters present. Beef tallow is a mixture of stearin, palmitin, and olein, with the more solid stearin predominating. Olive oil and cottonseed oil contain large proportions of olein and are therefore liquid.

299. Butter and Oleomargarine. Butter contains from 80 per cent to 85 per cent of a very complex fat consisting of several of the glycerin esters with olein and palmatin predominating. The characteristic flavor of the butter fat, however, is due to the presence of about 5 per cent of butyrin ($C_3H_5(C_4H_7O_2)_3$), which is the salt of butyric acid ($H\ C_4H_7O_2$). This fat occurs only in butter, and the methods of distinguishing true from adulterated butter depend upon this fact.

The high price of butter has led to the manufacture of certain butter substitutes. Oleomargarine is made from various animal fats, combined with cottonseed, peanut, and palm oils, the different fats being mixed in such proportions as to give a substance of about the same consistency as butter fat. To give the product a butter flavor, the melted fat is poured into ripened milk, that is, milk that has been soured as cream is in butter making, and is then churned. In butterine a certain proportion of butter fat is added to impart the butter flavor.

While oleomargarine is a valuable food product, the temptation to sell it for butter has proved so great that strict laws regulating its sale are required. Oleomargarine colored to resemble butter is taxed ten cents a pound by the United States government, and the sale of colored oleomargarine is entirely prohibited in some states. The fraudulent substitution of other fats for butter fat is the more reprehensible in view of the recent discovery that butter fat contains certain growth-producing substances, called vitamines, which are not found in lard and vegetable fats.

300. Renovated Butter. Much of the butter placed on the market becomes strong or rancid, and the " renovating " of this butter has become an important industry. The butter is melted and the curd and the water are removed as well as the scum on the top. Air is then forced through the melted fat until the disagreeable odors are removed and the fat is nearly tasteless. The fat is then churned with ripened milk as in the case of oleomargarine.

301. Making Solid Fats from Oils. A comparison of the formulas of the liquid fat olein and the solid fat stearin shows that the latter contains six more hydrogen atoms than olein. By treating olein with hydrogen in the presence of finely powdered nickel, which acts as a catalytic agent, it can be made to absorb hydrogen and thus be changed into stearin. This process is known as *hydrogenation of oils*. A number of edible lard substitutes are made by *hydrogenating* cottonseed oil until it has the consistency of lard.

Hydrogenation has proved of great value also in soap making. Oils that give soft soaps can be converted into compounds that yield the more valuable hard soaps. Fish oils, which have objectionable odors, can by this process be deodorized and made suitable for soap making.

302. Drying Oils. Linseed oil consists largely of the fat having the formula $C_3H_5(C_{18}H_{31}O_2)_3$. This fat is linolein, the glycerin salt of linoleic acid ($H \cdot C_{18}H_{31}O_2$). It will be noted that linolein has twelve fewer hydrogen atoms in the molecule than stearin. It has the property of absorbing oxygen from the air and thereby of being changed into a hard, solid substance. For this reason linseed oil is called a drying oil and it is this property that makes it valuable for use in paint making. A few other oils, for example, poppyseed oil and corn oil, have this property of drying in a less marked degree than linseed oil.

303. Soaps. All esters are decomposed when heated with the hydroxides of the alkalies. With stearin and sodium hydroxide the reaction is

$$\underset{\text{glycerin stearate}}{C_3H_5(C_{18}H_{35}O_2)_3} + 3\,NaOH \rightarrow \underset{\text{glycerin}}{C_3H_5(OH)_3} + \underset{\text{sodium stearate}}{3\,NaC_{18}H_{35}O_2}.$$

Sodium stearate is a soap. Reactions like the above may be made to take place between all fats and the hydroxides of sodium or potassium, and in each case a soap is formed. Soaps, therefore, may be said to be the *sodium or potassium salts of the fatty acids.* Sodium soaps are known as hard soaps, while those containing potassium are soft soaps.

In soap making the lye solution is gradually added to the oil or melted fat, which is kept warm and stirred by jets of steam. When the reaction appears to be complete, salt is added to the mixture and the soap separates and floats on top, as it is insoluble in a solution of salt. This process is called *salting out.* In this method of making soap most of the glycerin remains in the solution with the salt and spent lye and is hard to recover. In the old fashioned homemade soft soap (Fig. 137) the alkali used is potassium carbonate obtained by leaching wood ashes. The potassium is present in plants as the salts of organic acids, but all alkali salts of organic acids are changed to carbonates when burned:

$$K_2C_2O_4 + O \rightarrow K_2CO_3 + CO_2.$$

The larger soap factories first hydrolyze the fats with superheated steam, and a part of the liberated fatty acids are treated with sodium hydroxide to make the soap. The reaction between an alkali and a fat resulting in the production of a soap is known as *saponification.*

FIG. 137. — Making soft soap on the farm.

Calcium and magnesium salts of the fatty acids are insoluble in water, and these insoluble salts are formed when soap is added to hard water (119). As these insoluble compounds have no cleansing power, that part of the soap which reacts with the calcium and magnesium of hard water is wasted. The methods of determining hardness of water depend on the fact that a lather is not produced in water until all of the calcium and magnesium is precipitated (16).

304. Essential Oils. Many plants contain so-called essential or volatile oils which impart to them their characteristic taste or odor. These substances are not true oils and vary in character according to the source. They are completely volatilized when heated and leave no permanent greasy residue on cloth or paper. Oil of lemon, oil of peppermint, and oil of cedar are examples of essential oils. The odor of new-mown hay is due to a volatile oil. The clovers, particularly sweet clover, have characteristic essential oils. These oils are lost when the hay is overcured or exposed to leaching rains. Some of the essential oils of foods exert a favorable influence on digestion by imparting palatability to the food and stimulating the flow of the digestive fluids. Some of the essential oils have medicinal value, while others, as oil of bitter almonds, are poisonous.

EXERCISES

Ex. **173.** Name some of the common fats found in plants and animals. What is produced when fat is heated with steam under pressure? Do fats always consist of glycerin united with one or more fatty acids? What are the three most common fatty acids? How are the fats themselves named? Name the three most common fats. Do natural fats consist of one of these fats or of a mixture of two or more?

Ex. **174.** What gives the characteristic flavor to butter fat? Does this substance occur in any other natural fat? What is oleomargarine? How is the flavor of butter given to it? What is renovated butter?

Ex. **175.** Write out the formulas for olein and stearin. How many more hydrogen atoms are there in stearin than in olein? Can hydrogen be added to olein? What is this process called? Name a product that is so prepared. Of what value is hydrogenation in soap making?

Ex. **176.** (Teacher) Place 20 grams of lard in an evaporating dish or a granite-ware cup and warm to melt the lard. Dissolve 10 grams of sodium hydroxide in about 40 cc. of water. Add the solution slowly to the melted fat. Heat gently, with constant stirring, until a few drops of the mixture dissolve completely in clear water leaving no globules of fat. When the mixture is cool, add a strong solution of salt and the soap will separate and rise to the top, where it will finally solidify.

Ex. **177.** Explain the changes which take place in soap making. Dissolve some of the soap in water and add hydrochloric acid. The material which floats on the top consists of the fatty acids which were present in the lard and the soap. To another portion of the soap solution add a solution of calcium chloride. What is the curd in this case? Why is more soap necessary with hard water than with soft water?

Ex. **178.** (Teacher) Half fill a liter distilling flask with chopped green leaves of peppermint, spearmint, wild bergamot, or sweet clover, and add 200 cc. of water. Connect with a condenser and distill until the receiving flask contains about 50 cc. of liquid. A few drops of the essential oil will be found floating on the water.

CHAPTER XXXIII

CARBOHYDRATES

305. THE compounds that are produced most abundantly by growing plants are substances which contain carbon combined with hydrogen and oxygen, the latter two being present in the proportion in which they are found in water. Because of this relation between the hydrogen and oxygen these compounds are collectively known as *carbohydrates*. Their importance will be realized when it is said that this group includes the sugars, starches, woods, and all plant fibers.

306. Grape sugar, or glucose, as its name signifies, is the sugar found naturally in grapes. The whitish efflorescence on raisins and dried figs is glucose. When dry, glucose is a waxy mass that can be made to crystallize only with great difficulty. It is sweet to the taste, but its sweetening power is only about three fifths that of cane sugar. It occurs in many fruits. Commercially glucose appears as a heavy sirup and in this case is produced by the action of acids on starch (311). It is the principal constituent of many table sirups, especially those labeled corn sirup. It is also used in preserving fruits and in candy making, largely because the addition of about 10 per cent of glucose overcomes the tendency of cane sugar to crystallize.

Glucose is a good food and there is no reason for the popular prejudice against it, except that in the past it was used as an adulterant for the much sweeter cane sugar.

The formula for glucose is $C_6H_{12}O_6$. When acted upon by yeast it very readily ferments, forming ethyl alcohol and carbon dioxide (281). Glucose is a reducing agent (43) and the chemical test for it and some other sugars depends on this fact. The test reagent is known as Fehling's solution. It is made by mixing solutions of copper sulphate and sodium hydroxide and then adding Rochelle salt (292) until the precipitate first formed is redissolved. This liquid may now be considered as a solution of cupric oxide (CuO). If a reducing agent such as glucose is added to hot Fehling's solution, it abstracts oxygen from the cupric oxide, leaving cuprous oxide, which is precipitated:

$$2\,CuO \rightarrow Cu_2O + O.$$

307. Fruit sugar, or fructose, is a sugar found in some fruits. It occurs also in the nectar of flowers and, therefore, is a constituent of honey, in which it is associated with glucose. It is sweeter than glucose. It ferments under the action of yeast but not so readily as glucose.

The formula for fructose is $C_6H_{12}O_6$, which will be seen to be the same as that assigned to glucose. It often happens in organic chemistry that two or more entirely distinct compounds have the same kind and number of atoms in the molecule. The explanation offered is that the properties of a compound depend not solely upon the number and kind of atoms in the molecule, but upon the way in which these atoms are arranged. A child may with the same number of red, white, and blue blocks work out a variety of patterns by different arrangements of the blocks. Likewise the same number of carbon, hydrogen, and oxygen atoms may be arranged in the molecule in a number of ways, and each different arrangement results in a distinct compound with some properties peculiar to itself. Two or more compounds that have the same general formula are known as *isomeric compounds*, or *isomers*.

308. Cane Sugar. As ordinarily used the term sugar refers to cane sugar, or sucrose, which has the formula $C_{12}H_{22}O_{11}$. This sugar is found in sugar cane (Fig. 138), sorghum, sugar beet, and the sap of the maple tree, as well as in many fruits.

FIG. 138. — Growing sugar cane.

There is more of this sugar in nature than any other sugar. About half the sugar of commerce comes from sugar beets, the rest largely from sugar cane. The juice of the cane, or the water extract of the beet, is first treated with milk of lime (114) to neutralize the acids and to precipitate the albuminous substances that are present. The excess of lime is removed by passing carbon dioxide through the liquid, and the clarified liquid is evaporated in vacuum pans until very thick. Upon cooling, the sugar separates and the remaining liquid is removed by whirling in a centrifugal machine. The product is a brown sugar which is refined by dissolving it in water and filtering the solution through bone black (91) to remove the coloring matter, and again concentrating in vacuum pans. This treatment produces the clear crystals sold under the name of granulated sugar. The liquid left after the separation of the brown sugar is called molasses. The annual production of sugar from cane and sugar beets amounts to over ten million tons. Whether produced from cane or beets the sugar is the same compound. The popular notion that cane sugar differs from beet sugar is erroneous.

Cane sugar when pure does not reduce Fehling's solution. If the sugar is boiled with dilute acids (sulphuric, for instance) and the acid is neutralized with an alkali, the resulting solution has a strong reducing action. This is due to the fact that the molecule of cane sugar takes on water and is changed into a molecule each of glucose and fructose:

$$\underset{\text{cane sugar}}{C_{12}H_{22}O_{11}} + H_2O \rightarrow \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}.$$

The acid apparently acts as a catalytic agent. This change of cane sugar into glucose and fructose is called inversion, and the product (the mixture of glucose and fructose) is known as *invert sugar*. In jelly making the acid of the fruit always inverts a part of the sugar. Vinegar is sometimes added in candy making to invert part of the sugar so that the glucose and fructose will check the tendency of the cane sugar to crystallize.

Cane sugar does not ferment readily. After a time, however, it does undergo alcoholic fermentation, for the yeast contains a substance which slowly inverts the cane sugar, and the resulting glucose and fructose readily ferment.

309. Malt sugar, or maltose, is a sugar having the same formula as cane sugar (isomer) and is formed by the action of malt on starch (311). Malt, which is made by steeping barley in water until it germinates, and then drying it, contains a substance called diastase, which has the power of changing starch into maltose ($C_{12}H_{22}O_{11}$). Maltose is a reducing sugar. It is easily fermented and is the intermediate product in the formation of alcohol from starch. When heated with dilute acid it is changed into glucose:

$$\underset{\text{maltose}}{C_{12}H_{22}O_{11}} + H_2O \rightarrow 2\ \underset{\text{glucose}}{C_6H_{12}O_6}.$$

310. Milk sugar, or lactose, is another isomer of cane sugar that is found only in milk. It is produced commercially from the whey left in cheese making. It is much less sweet and less soluble than cane sugar. It is the sugar most commonly used in medicine. It is a reducing sugar but does not readily ferment. Its most important reaction is the change into lactic acid (290):

$$\underset{\text{lactose}}{C_{12}H_{22}O_{11}} + H_2O \rightarrow 4\ \underset{\text{lactic acid}}{H\cdot C_3H_5O_3}.$$

311. Starch is one of the most abundant compounds produced by plants. It forms from 50 per cent to 75 per cent of the dry matter of seeds and tubers, where it is stored as food for the young plant before it is able to obtain its own food. In this country most of the starch of commerce is prepared from corn, while in Europe potatoes are the principal source.

Starch is separated from corn by soaking, grinding, and washing the grain in water and then filtering. In the last process, the finely divided starch passes through bolting cloth and then is allowed to settle from the water in which it is suspended. Starch occurs in the plant as granules, which vary in form and markings according to the plant from which they are derived (Fig. 139). It is possible to tell the source of the starch by examination under the microscope.

Starch is not soluble in water. When treated with boiling water the granules swell and burst, forming a gelatinous mass known as starch paste. It is colored blue by iodine. It has the composition represented by the formula $C_6H_{10}O_5$; but there are good reasons for believing that the molecule is really much larger than this formula suggests. It is quite customary, therefore, to write the formula $(C_6H_{10}O_5)_x$, in which x represents an unknown number. Starch is changed into glucose by prolonged boiling with dilute acids (306):

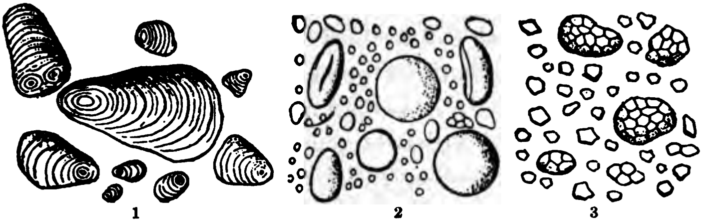

Fig. 139. — Magnified starch granules. 1. Potato. 2. Wheat. 3. Rice.

$$C_6H_{10}O_5 + H_2O \rightarrow C_6H_{12}O_6.$$

Large quantities of starch are used for the production of glucose, which is prepared commercially by boiling corn starch under pressure with hydrochloric acid. The acid is then neutralized with sodium carbonate, and after the liquid has been clarified with bone black it is concentrated to a thick sirup containing 30 to 40 per cent of glucose, mixed with dextrin (312).

Starch is a valuable constituent of many foods. It is used in the laundry, for the making of library or photo pastes, in the finishing of cotton cloth, and for many other purposes. Sago is a starch prepared from certain palm trees. Tapioca, a starch, comes from the root of a tropical plant called cassava.

312. Dextrin. When starch is heated to about 200° C. it is changed into a pale yellow powder that is soluble in water and is not colored by iodine. This substance is called dextrin and is represented by the same general formula as starch $(C_6H_{10}O_5)_x$. Its solution is gummy, and dextrin is now used as a substitute for the natural gums in making mucilage. Dextrin is an intermediate product in the formation of glucose from starch; consequently commercial glucose usually contains dextrin as well as glucose. In laundry work the heat of the iron converts some of the starch into dextrin, which gives a glossy finish to the cloth.

313. Cellulose is the substance that forms the basis of the woody fiber of plants. It is found most abundantly in the stems, roots, and leaves of plants, particularly at maturity. It is the structural basis of the vegetable world and forms the framework of every plant cell. In some plants it is the most abundant material present; in hay and coarse fodder it makes up 30 to 40 per cent of the dry matter. Cotton and linen are examples of almost pure cellulose. When quite pure, cellulose is a colorless material insoluble in water and differing in texture according to its source. It is assigned the same general formula as starch $(C_6H_{10}O_5)_x$. In the case of cellulose, however, x probably stands for a larger multiple than it does in the formula for starch. It is evident that cellulose plays an important rôle in providing material for fuel, shelter, and raiment for mankind. It is used in paper making and in the preparation of guncotton, celluloid, and many other useful materials.

314. Nitrogen Compounds of Cellulose. Nitric acid unites with cellulose in several proportions, and the resulting compounds are known collectively as *nitrocellulose,* or *guncotton.* This substance, which is highly explosive, is the basis of *smokeless powder,* which has supplanted black gunpowder in military operations. Some of the nitrocellulose compounds dissolve in a mixture of alcohol and ether, forming the substance known as *collodion.* Collodion is used as liquid court-plaster and in the making of photographic films, because on the evaporation of the ether and alcohol a tough, transparent film remains. *Celluloid* is made by combining guncotton with gum camphor. It is, therefore, very inflammable, and many serious accidents have occurred from carelessly allowing celluloid to come into contact with a flame. A mixture of guncotton and nitroglycerin is used under the name of *blasting gelatin,* especially for heavy blasting, as it is a very powerful explosive.

315. Gums and Pectin Bodies. Closely related to the sugars are the gums, such as gum arabic and those that exude from peach and cherry trees. Gum arabic is an isomer of cane sugar ($C_{12}H_{22}O_{11}$). It is used in making mucilage and in the better grades of gumdrops. Pectin bodies are found in many fruits and some vegetables. They are jellylike substances that are soluble in hot water and are commonly known as fruit jellies. Gums and the pectin bodies are considered to have the same food value as starch and sugar.

EXERCISES

Ex. **179.** What is meant by the term carbohydrate? Are the carbohydrates of much importance? What is the formula for glucose? Where is it found in nature? How is commercial glucose manufactured? For what is it used?

Ex. **180**. Heat a little Fehling's solution to boiling and add a few drops of a solution of pure glucose, or of Karo sirup. Heat again. The red precipitate of cuprous oxide is the test for glucose or other reducing sugar. Explain the action of glucose on Fehling's solution. Soak a few raisins in water and test the solution for grape sugar.

Ex. **181**. What is the formula for fructose? Compare with that for glucose. How do you explain the fact that there are two compounds with the same formula? What is an isomer? What sugars are found in honey? Test a sample of honey with Fehling's solution.

Ex. **182**. What sugar is found most abundantly in nature? Name some of the plants that contain it. From what plants is it prepared on a commercial scale? How is the brown sugar changed to white sugar? What is molasses? Is there any difference between sugar from cane and from beets? What is the formula for cane sugar?

Ex. **183**. Test a solution of granulated sugar with Fehling's solution. To a little of the solution add a few drops of hydrochloric acid and boil two minutes. Add a pinch of sodium carbonate to neutralize the acid and test with Fehling's solution. What change has taken place in the sugar?

Ex. **184**. Test a sample of jelly from home for glucose. Fruit juice and cane sugar were used in making the jelly — account for the presence of glucose. Why is vinegar or cream of tartar sometimes used in homemade candies?

Ex. **185**. What is the formula for maltose? How is it prepared? What is the formula for milk sugar? Compare this with the formula for cane sugar. Are they isomers? What is the sole source of milk sugar? What product is formed from milk sugar when milk sours?

Ex. **186**. Reduce two large potatoes to a pulp with a vegetable grater. Tie the pulp in a clean thin cloth and squeeze it into a vessel of water, occasionally dipping the bag into the water. Pour the liquid into a tall cylinder and allow the starch to settle. Pour off the clear water, and transfer the starch to a shallow dish and allow it to dry.

Ex. **187**. What can you say about the importance of starch? From what is the starch of commerce produced? Is starch soluble in water? What happens to it when boiled with water? What is the formula? Why is the formula multiplied by x? Give some important uses for starch.

Ex. 188. Put a little of the starch from Ex. 186 into a test tube and boil it with water. Add a drop of a solution of iodine and potassium iodide. What change takes place? This is the test for starch. Test for the presence of starch in various seeds and vegetables.

Ex. 189. Boil a little starch in water and test part of the liquid with Fehling's solution. What is the result? To another part of the boiled starch add a few drops of hydrochloric acid and boil three or four minutes. Test with Fehling's solution. What difference do you note? What change has taken place in the starch? How is commercial glucose sirup prepared?

Ex. 190. How is dextrin prepared? What is the formula? Is it soluble in water? Is it colored by iodine? For what is it used? Why does the crust of bread have a sweeter taste than the crumb?

Ex. 191. (Teacher) If a compound microscope is at hand prepare slides of starch from potatoes, corn, and other food products. Have the class observe and make drawings of the different starch grains.

Ex. 192. What is cellulose? In what parts of the plant is it most abundant? Is it an important substance? Give two examples of almost pure cellulose. What is its formula? What explosive materials are prepared from cellulose? What is collodion? Celluloid?

Ex. 193. What are the substances known as gums? Give the formula of gum arabic. Compare this with the formula for cane sugar. For what is it used? What are the pectins and where are they found? Of what importance is pectin in jelly making?

CHAPTER XXXIV

ORGANIC NITROGEN COMPOUNDS

316. THE organic compounds so far studied, with the exception of nitroglycerin and nitrocellulose, contain not to exceed three elements; namely, carbon, hydrogen, and oxygen. If a piece of lean meat or a bit of dried egg white is placed in a deep test tube and heated, the odor of ammonia can be detected in the escaping gases. A piece of moistened red litmus paper held in these gases will be changed to blue. This formation of ammonia from the meat or egg will take place more rapidly if the material is first mixed with soda-lime. Evidently, then, these substances contain nitrogen, or the ammonia could not have been formed. During the heating, water vapor is given off, and charcoal remains in the tube; the meat and egg, therefore, must contain carbon, hydrogen, and oxygen as well as nitrogen. If a piece of paper moistened with sugar of lead is held in the escaping gases in the above experiment, it will be blackened, showing the presence of hydrogen sulphide, which indicates that sulphur also is found in the egg and meat. Both meat and egg belong to a large class of compounds known as proteins.

317. Proteins are generally distributed in the animal and vegetable kingdoms. They are not nearly so abundant as the carbohydrates. It is estimated that ten times as much carbohydrate as protein is produced in the vegetable kingdom. All proteins contain the four elements, carbon, oxygen, hydrogen, and nitrogen, and most of them contain sulphur as well. Some of the proteins contain in addition to the five elements mentioned, a small amount of phosphorus. No one has been able to determine the formula for any of the proteins, and they are thought to be the most complex of all the chemical compounds. The complexity of the molecule may be inferred from the fact that the most careful experimenters estimate the molecular weight anywhere from 16,000 to 50,000. When this molecular weight is compared with that for sugar ($C_{12}H_{22}O_{11}$), which is 342, some idea may be formed of what an enormous number of atoms the protein molecule must contain. Most of the proteins found in nature are insoluble in water, although there are a number of soluble proteins, the white of egg, for example. Some which will not dissolve in water are soluble in weak salt solutions. The proteins unite with dilute acids and alkalies, forming compounds that are sometimes soluble in water and sometimes insoluble. Concentrated acids and alkalies dissolve all the proteins. The proteins decay very readily, and their decomposition is accompanied by offensive odors, as for example that of rotten eggs.

318. Tests for Proteins. While there are a great many kinds and classes of proteins, there are a few tests that apply to all proteins, which may be used to show the presence of protein in any substance that is being examined:

(1) Nitric acid gives a permanent yellow color with proteins upon warming. The yellow color produced when nitric acid is dropped on the hands is due to the reaction of the nitric acid with the protein of the skin. If ammonia water is added to the protein which was turned yellow by nitric acid, it will be changed to an orange color.

(2) A solution of mercury in nitric acid (mercuric nitrate), known as Millon's reagent, gives a brick red color with proteins when heated.

(3) Tannic acid forms a tough, leathery compound with proteins. If the protein is in solution, tannic acid throws it down in the form of a leathery precipitate. The change of animal hides into leather is largely due to the formation of tough, insoluble compounds by the action of the tannin on the proteins of the skin or hide.

319. Proteins Insoluble in Water. The muscle, or lean meat, of all animals consists largely of insoluble protein, although there is a small amount of soluble protein present, as will be seen later. Another important source of insoluble proteins is the wheat kernel. If a cupful of wheat flour is made into dough with a small quantity of water and then wrapped in a thin cloth and kneaded under water (Fig. 140), the starch of the flour will pass through the cloth, leaving a sticky, elastic mass (Fig. 141), which is the protein material known as the *gluten* of the wheat. Gluten consists of two proteins; one, *gliadin*, is a glue-like body that binds together the particles of flour in the dough; the other, *glutenin*, is a fine, gray material that does not have the binding property of the gliadin. It is the gliadin in wheat that makes possible a light loaf of bread. All the grains contain proteins similar to the gluten of wheat, although rye is the only other grain in which the gluten is of the quality to make a good dough.

FIG. 140. — The separation of gluten from wheat.

FIG. 141. — Showing the elasticity of gluten.

Casein, the important protein of milk, seems to be dissolved but is really held in mechanical suspension. Its important characteristic is its behavior toward acids and rennet. When either an acid or rennet is added to milk, the casein separates in a thick curd. Rennet extract is made by soaking the linings of the stomachs of young calves in a solution of common salt. The young stomach contains a substance called rennin which has the property of coagulating casein. When milk sours naturally, the lactic acid unites with the casein and forms a curd. The curds formed by acids and that by rennet are both used in cheese making.

320. Albumins are proteins that are soluble in water. The most familiar example is the white of egg. Albumins coagulate when heated, as is well shown when an egg is boiled or fried until the white becomes hard. If a solution of egg white is boiled, the albumin coagulates and is precipitated. Albumin occurs to a limited extent in meat, as can be shown by rubbing chopped raw meat in a mortar with water, filtering it, and then heating the filtrate. When meat is boiled a scum of albumin is often found on the surface of the water, especially if the meat was placed in cold water in the beginning. Albumin is present in milk and is left in solution in the whey when casein is coagulated either by acid or rennet. Albumin is also one of the constituents of the blood.

321. Peptones are proteins that are soluble in water and are not coagulated by heat. They are usually formed by the actions of ferments on the other proteins. When gastric juice, for instance, acts on meat or coagulated albumin, the protein disappears, forming a more or less clear solution. Since this solution gives no precipitate when heated, it does not contain albumin; but by applying the tests described above (318) it can be shown that protein is present. Pepsin, a substance prepared from the stomach of the pig, can be used to perform this experiment in the laboratory. Some of the so-called predigested or peptonized foods contain peptones prepared by the artificial digestion of the food by means of pepsin. The proteins of the food must be converted into peptones in the stomach and intestines before they can be absorbed and pass into the blood. Peptones are not present in ordinary foods in appreciable amounts but are formed from the foods during digestion.

322. Importance of Proteins. Proteins are present in all plant and animal cells. The vital part of the cell, the protoplasm, consists of protein material, and consequently it will be seen that all life depends on the proteins. The muscles of animals consist largely of proteins; and the only way in which the muscles can be built up and repaired is by means of the proteins of the food, since the animal body is not able to manufacture proteins from other materials. The proteins, then, have a place that cannot be filled by the carbohydrates, fats, or any other compounds. All parts of the plant contain some protein; but the seeds contain the most, since it is stored there for the use of the plantlet when it begins its growth. It will be seen later that the plant can manufacture these complex proteins from very simple substances.

323. Gelatin and Other Albuminoids. There is another class of nitrogen compounds, somewhat resembling the proteins, to which the name albuminoids has been given. The best-known example of this class is gelatin, which is obtained from the connective tissue and bones of animals. Commercial glue is an impure gelatin. Gelatin dissolves in hot water and forms a jellylike mass upon cooling. While it is a good food product and is easily digested, it will not take the place of protein in the food. Keratin is a hard, horny albuminoid found in the horns, hoofs, hair, nails, and feathers. Mucin, the chief constituent of mucus, gives the sliminess to the secretion of the mucous membrane. It is present in the saliva.

324. Amines. There is a marked difference between the simple nitrogen compounds that the plants take in from the soil (nitrates and ammonia) and the very complex proteins that they store in their seeds. It is not surprising, therefore, to find that there are intermediate compounds, or, in other words, that the building up of a protein is not a matter of one change but takes place by steps. Some of the intermediate compounds are known and have been assigned the name of amines. Their chief interest for the purpose of this text lies in the fact that they are intermediate compounds in the building up of proteins, and that when the protein is broken down during digestion or decay, the amines are produced before the nitrogen is finally changed back to ammonia or nitric acid (168).

325. Protein Production and Destruction. The plant produces protein by a series of changes about as follows:

(1) Ammonia or nitrate is taken from the soil.
(2) An amine is formed from ammonia.
(3) A protein is finally formed from the amine.

When animals consume the plant as food the reverse order of changes takes place:

(1) The protein is digested and made over into animal protein.

(2) The animal protein is finally broken down into amines.

(3) The amine is expelled from the body as waste matter.

(4) In the soil the amine is changed into ammonia and nitric acid by the bacteria and is ready to begin again this cycle of changes.

326. Alkaloids are nitrogenous organic compounds found in many animals and plants, but not to any appreciable amount in true food plants. They somewhat resemble ammonia in their chemical behavior and are, therefore, called alkaloids (*like alkalies*). They form the so-called *active principle* of many of the plants used in medicine. The following are some of the principal alkaloids:

Caffeine ($C_8H_{10}N_4O_2$), from coffee and tea.
Quinine ($C_{20}H_{24}N_2O_2$), from Peruvian bark.
Strychnine ($C_{21}H_{22}N_2O_2$), from nux vomica bean.
Morphine ($C_{17}H_{19}NO_3$), from the poppy.
Nicotine ($C_{10}H_{14}N_2$), from tobacco.

Alkaloids are also formed at times in the animal body. Some of the alkaloids formed during the decomposition of proteins are extremely poisonous and are known as ptomaines. They sometimes occur in stale meat, fish, and cheese.

EXERCISES

Ex. **194.** Place a piece of lean meat in a test tube fitted with a cork in which is a small glass tube. Heat over the flame. Note the odor of the escaping gas. Hold a piece of moist red litmus paper in the gas. Test with sugar of lead paper. Repeat the experiment with some peas or beans. Mix a quarter of a test tube full of dry clover with an equal bulk of soda-lime and heat and test as above. Have you any evidence that ammonia was formed? Was sulphur present in the gas? What was the source of these substances?

Ex. **195.** What can you say about the distribution of proteins in nature? Are they as abundant as the carbohydrates? What chemical elements are found in the proteins? Is the protein molecule very complex? What effect do acids and alkalies have on proteins?

Ex. **196.** To a little egg white add a few drops of Millon's reagent and warm the mixture. What is the result? Try the same experiment with crushed wheat or corn, milk, or other foodstuffs. Which gave the test for proteins?

Ex. **197.** (Teacher) Mix a cupful of flour to a dough with water. Place it in a cloth bag and knead it under water (Fig. 140) until all the starch has been washed out. Examine the material left in the cloth. Is it sticky and elastic? What is this substance? Make the test with Millon's reagent on a small piece of it. What other substances can you name that belong to the group of insoluble proteins?

Ex. **198.** Dissolve the white of a raw egg in about a pint of water and use it for the following tests: (1) Heat a part of the solution to boiling in a test tube. What happens? (2) To another portion of the egg-white solution add tannic acid. What is the result?

Ex. **199.** Place 10 grams of ground oats in a bottle with 50 cc. of water. Cork and shake the bottle vigorously and let it stand for half an hour or until the next period. Filter and test the filtrate with tannic acid. Is there any albumin in the oats? Mention some other materials which contain albumin.

Ex. **200.** (Teacher) Dissolve five grams of commercial scale pepsin in a quart of water containing 5 drops of hydrochloric acid. Pass the white of a hard-boiled egg through a sieve and place it in a flask with 250 cc. of the above pepsin solution. Place the flask in a water bath and keep at blood temperature for four or five hours. What action did the pepsin have on the congested egg albumin? Filter some of the solution for use in the next exercise.

Ex. **201.** Heat a portion of the solution from the last exercise to ascertain whether it contains albumin. To another portion add tannic acid. To another add Millon's reagent and heat it. Does the solution contain albumin? Does it contain some kind of protein? What are these soluble proteins called? Are peptones present in large quantities in foods? Are they produced during digestion?

Ex. **202**. Why are proteins vital to plant and animal life? Where does the animal get the material to build its muscle? Can the animal manufacture proteins? In what part of the plant is the most protein found? Why is the protein stored in the seed? Can the plants manufacture protein?

Ex. **203**. (Teacher) Obtain a package of any of the culinary gelatins and prepare a jelly with the amount of water recommended in the printed directions. To what class of compounds does gelatin belong? What is the source of gelatin? What is the difference between gelatin and glue? Name an albuminoid found in bones and hair; one found in mucus.

Ex. **204**. Does the plant form intermediate compounds between the nitrates and the proteins? What are some of the compounds called? Outline the cycle of changes in the production and destruction of proteins. What are the alkaloids? Name some of the common alkaloids. What is meant by a ptomaine?

CHAPTER XXXV

COMPOSITION OF PLANTS

327. Water. All plants and, indeed, all food materials contain water which in many cases makes up the larger portion of their weight. Green plants, such as the mature corn plant, contain as much as 80 per cent of water, while some of the more succulent plants like cabbages, lettuce, and spinach may contain as high as 90 per cent of water. In general it may be said that the younger the plant the larger the percentage of water it contains, and that as the plant matures the percentage of water is decreased. The stems or woody parts of plants contain less water than the leaves. The juicy fruits, such as oranges, and strawberries, often contain over 90 per cent of water, while the grains may be as low as 10–15 per cent in moisture content. Many of the plants and plant products are partially dried to remove the water before being used as foods for man or domestic animals, but even well-cured hay, or such a dry substance as wheat flour, still contains about 10 per cent of water.

The amount of moisture in a substance is determined in the laboratory by heating the material for several hours at the temperature of boiling water. By this means the water in the plant or other substance is converted into steam and thus expelled. The material is usually dried in a water oven similar to the one shown in Fig. 142. The walls of this oven are double, the space between them being partially filled with water which is kept boiling by means of a gas burner placed beneath the oven. In such an oven the substance being dried may be kept heated without danger of the temperature rising above 100° C. and scorching the material. Ovens heated by electricity are also used, the heat being adjusted by a regulator which prevents too great a rise in temperature.

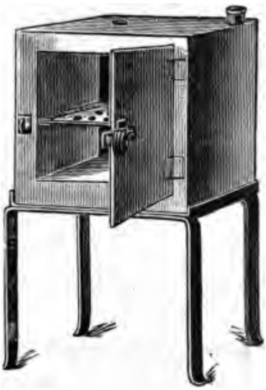

FIG. 142. — Water oven for the determination of moisture in plant and animal products.

328. Dry Matter. The dry matter of a material is the portion left after all the water has been removed. The dry matter in plant products varies within wide limits, being as low as 5–6 per cent in some fruits and over 90 per cent in certain cereal products. The amount and composition of the dry matter in a food material determines its value as a nutrient.

329. Plant Ash. The ash of a plant or of any substance is that portion that remains after the substance is burned at the lowest temperature necessary for complete combustion. It corresponds in a general way to the ashes left in the stove when wood is burned. The ash of the plant is sometimes spoken of as the *mineral matter,* or as the *inorganic matter* of the plant, and also as the non-volatile part. It includes all of the material which the plant obtained from the soil with the exception of the water and nitrogen, and possibly part of the sulphur and phosphorus, which disappear when the plant is burned.

The percentage of ash is determined in the laboratory by igniting a small quantity of the material (two grams) in a platinum or porcelain dish (Fig. 17) until all the carbon is burned and a white or nearly white ash remains. During the experiment the material must be carefully protected from currents of air, which might blow away the light particles of the ash. The weight of the ash obtained in this way divided by the weight of the material taken for the experiment gives the percentage of ash in the substance.

330. Composition and Amount of Ash in Plants. The ash of all plants contains measurable quantities of nine chemical elements — potassium, calcium, sodium, iron, magnesium, phosphorus, sulphur, chlorine, and silicon. Traces of aluminum and manganese also are found in the ash. None of the elements of the ash existed in the plant in the elementary or free state; but they are always in chemical combination forming salts, or are combined with the elements that form the organic part of the plant. It has been shown, for example, that potassium occurs in grapes as acid potassium tartrate (292) and that phosphorus and sulphur are contained in certain of the proteins (317).

Some plants contain much more ash than others, and there is great variation in the percentage composition of the ash of different plants. The ash of clover hay contains nearly four times as much calcium as the ash of timothy hay, but only one fifth as much silicon. The ash is not evenly distributed throughout all parts of the plant. In corn the amount of ash in the different parts of the plant is as follows:

Roots	5.8 per cent
Leaves	8.1 per cent
Stems	6.6 per cent
Grain	1.4 per cent

331. Organic Matter. That part of the plant which completely burns and passes off in the form of gaseous products is termed organic matter. It is determined in the laboratory by subtracting the percentage of ash from the dry matter. The organic matter includes the proteins, carbohydrates, fats, organic acids, and other so-called organic compounds of the plant. It is customary to divide the organic constituents of the plant into two large classes; namely, (1) nitrogenous, and (2) non-nitrogenous, the division depending upon the presence or absence of the element nitrogen in the compounds.

332. Nitrogenous Compounds. The proteins are the most important of the nitrogenous constituents of plants. Most of the proteins found in plants belong to the insoluble class, although the albumins are found to a limited extent. There are doubtless a great number of different plant proteins, but only a few of them have been isolated in the pure state and carefully studied. The proteins are found in all parts of the plant, but are much more abundant in the seeds. The leaves contain more protein than do the stems or roots.

All parts of a plant are composed of minute cells, lying close together. These cells are so small that the compound microscope must be used in order to see them. The cells contain a clear, granular substance, called protoplasm, which has about the consistency of the white of egg. It is the living substance of the cell, and the growth and functions of the plant depend upon the activity of the protoplasm. It is not known exactly what protoplasm is, but it is undoubtedly largely composed of proteins.

To find the amount of protein in a food material the chemist first determines the percentage of nitrogen and then multiplies this by 6.25 to obtain the equivalent of protein. This is because proteins contain, on the average, about 16 per cent nitrogen, or there is about one part of nitrogen to every 6.25 parts of protein (100 ÷ 16 = 6.25). The method of determining the nitrogen consists in first heating a weighed quantity of the food with sulphuric acid, which converts all the nitrogen into ammonium sulphate. The ammonia from the ammonium sulphate is then liberated by means of sodium hydroxide (167) and its amount determined. The chemist calls the material determined by multiplying the nitrogen by 6.25 *crude protein,* because it is not quite true that all the nitrogen in the food is in the form of protein.

333. The non-nitrogenous compounds of plants are (*a*) fats; (*b*) cellulose, or fiber; (*c*) carbohydrates.

(*a*) *Fats,* or oils, are present to some extent in all parts of the plant, but by far the larger portion is found in the seeds. The plant fats are for the most part oils, and the amount present varies greatly in different plants and in different parts of the same plant. Wheat contains about 2 per cent of fat; corn, 5 per cent; and flaxseed, 35 per cent or more of oil. The roots and stems of corn contain only about one half of one per cent of oil or fat. To determine the amount of fat, a weighed sample of food material is repeatedly washed with ether until all the fat is extracted; the ether is then evaporated and the fat weighed. In practice a special apparatus (Fig. 143) is used in which the same quantity of ether is made to pass

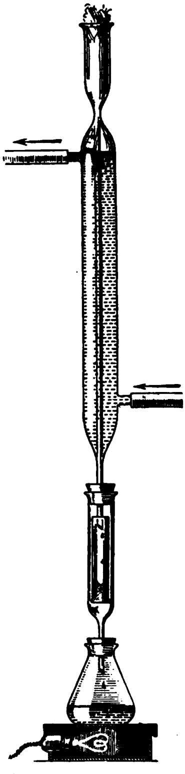

FIG. 143. — **Apparatus used in determining amount of fat or ether extract.**

repeatedly through the food material until all the fat is removed. The substance extracted from the food in this way is not pure fat, or oil, as the ether will also dissolve chlorophyll, waxes, and resins if they are present. In the case of the materials ordinarily used as food for man and animals the amounts of these substances present are so small as to cause no serious error in the determination; but to avoid inaccuracy of statement the substance determined as above is quite commonly called *ether extract* or *crude fat.*

(*b*) *Cellulose.* Cellulose, which forms the walls of the plant cells, and the closely related woody substances found in plants are commonly called fiber, or crude fiber. The determination of the amount of crude fiber in a food consists of three steps: (1) the fat is extracted from the substance with ether; (2) the material is then boiled with very dilute sulphuric acid to convert the starch into sugar, which is then washed away; (3) it is then boiled with dilute sodium hydroxide to dissolve the proteins (317), and the remaining material is washed, dried, and weighed as crude fiber. The roots and stems of plants contain the largest amounts of fiber, as may be surmised from their woody nature. The seeds usually contain very little fiber, only 2 to 5 per cent.

(*c*) *Carbohydrates.* While cellulose is strictly speaking a carbohydrate, this term in food analysis is commonly used to designate the starches and sugars, the cellulose being stated separately as fiber. Starch is the most abundant of the carbohydrates and is found principally in the seeds, roots, and tubers of plants, being stored in those parts which are concerned with new growth. Although the sugars are not so abundant in nature as starch, they are quite widely distributed and are found in large amounts in some fruits and vegetables and in the saps of sugar cane and the maple tree.

The separation and determination of the individual carbohydrates is a tedious process, and consequently, since they all have about the same food value, they are usually lumped together in the statement of analysis. In reality the carbohydrates are not determined at all, but the percentages of water, ash, crude protein, ether extract, and crude fiber are added together; their sum is then subtracted from 100, and the remainder is the percentage of carbohydrates. Wheat, for example, contains:

Water	9.25 per cent
Ash	2.95 per cent
Crude protein	13.25 per cent
Ether extract	2.20 per cent
Crude fiber	2.25 per cent
Total	29.90 per cent

100 − 29.90 = 70.10 = per cent of carbohydrates.

In the statement of analysis of a feeding stuff for domestic animals the expression *nitrogen-free extract* is used instead of carbohydrates. In the analysis of human foods the latter term is commonly used. In foods or feeding stuffs of vegetable origin the amount of nitrogen-free extract, or carbohydrates, usually exceeds that of any other of the groups of compounds.

334. Analysis of the Corn Plant. If a number of samples of the corn plant are taken for analysis as a feeding stuff at the time the corn is just mature, the average composition is found to be about as follows:

Water	79.3 per cent
Ash	1.2 per cent
Crude protein	1.8 per cent
Ether extract	0.5 per cent
Crude fiber	5.0 per cent
Nitrogen-free extract	12.2 per cent
Total	100.0 per cent

The following table states the composition more in detail. The numbers show the percentage of the various constituents.

Corn Plant 100	Water 79.3	Hydrogen 8.81				
		Oxygen 70.49				
	Dry matter 20.7	Organic matter 19.5	Crude protein	1.8	Nitrogen	0.29
			Ether extract	0.5	Carbon	9.05
			Crude fiber .	5.0	Oxygen	8.89
			Nitrogen-free extract . .	12.2	Hydrogen	1.27
		Ash . . 1.2	Chlorine . .	0.04		
			Potassium .	0.33		
			Phosphorus .	0.05		
			Calcium . .	0.12		
			Magnesium .	0.09		
			Iron . . .	0.02		
			Sulphur . .	0.01		
			Sodium . .	0.03		
			Silicon . . .	0.11		
			Oxygen . .	0.40		

It is worth noting that this table shows that over 98½ per cent of the green plant is composed of three elements, carbon, hydrogen, and oxygen, while all other elements combined make up only about 1½ per cent of its weight.

EXERCISES

Ex. **205**. Recall an experiment that shows the presence of water in plants. About how much water do growing plants contain? Which contain the larger percentage of water, old or young plants? Is there any water in cured hay and grains? How is the amount of water in a plant determined in the laboratory? (Note to the teacher. If a good balance and water oven are available the student should make a determination of the amount of water in a green plant.)

Ex. **206**. Burn some plant material in a porcelain dish. Stir the material until the ash is nearly white. What is included in the ash of the plant? How is the exact amount of ash in the plant determined? What is meant by the dry matter of the plant? Do the chemical elements exist in the plant in the same compounds in which they are found in the ash? Are the ash elements evenly distributed in all parts of the plant? Which part contains the least ash? The most ash?

Ex. **207.** What is meant by the organic matter of the plant? How does it differ from dry matter? Recall an experiment to show the presence of proteins in plants. Which class of proteins is present in plants in largest quantity? In what part of the plant are the proteins most abundant? What is protoplasm? Why are proteins so important to the plant? How is crude protein determined in the laboratory? Why is the nitrogen multiplied by 6.25?

Ex. **208.** Place a tablespoonful of corn meal or ground flaxseed in a small bottle and add an ounce of ether. Cork the bottle and shake it at intervals for an hour. Filter the contents through a dry filter into a glass dish and place it in the open window until the ether evaporates. (Caution. Do not handle the ether near a flame.) Did the ether extract any fat from the meal? How is fat in foods determined in the laboratory? Which part of the plant usually contains the most fat?

Ex. **209.** Place about one gram of ground straw or hay in a beaker and add 200 cc. of water and 20 drops of sulphuric acid. Boil on a sand bath 20 minutes. Allow the material to settle and pour off the liquid. Add 100 cc. of water and when the material again settles pour off as before. Now add 200 cc. of water and 4 cc. of the sodium hydroxide solution of the laboratory (10 %), boil 20 minutes, and wash as before. The material left in the beaker is crude fiber.

Ex. **210.** What is the crude fiber of the plant? How is the percentage of crude fiber determined? What parts of the plant contain the most fiber? Are the other carbohydrates of the plant determined directly? How are they calculated? What does "nitrogen-free extract" mean? Discuss the composition of the corn plant.

CHAPTER XXXVI

CHEMISTRY OF PLANT GROWTH

335. Seeds. Most of the crops grown by farmers and gardeners are raised from seeds. A careful examination of a seed shows that it consists of an embryo plant (Fig. 144 *pl*) surrounded by reserve food materials in the form of mineral matter and nitrogenous and non-nitrogenous organic compounds. An analysis would show that while the total amount of ash in the seed is small it is especially high in phosphorus, potassium, and magnesium — ash elements that are of great importance in the nutrition of the young plant. The non-nitrogenous part consists in most seeds largely of starch, with some oil and cellulose, and a little sugar and gums. In such seeds as flax, rape, mustard, and cotton seed, oil is the principal non-nitrogenous substance. Oil seeds are, as a rule, small in size but concentrated in food materials. The nitrogenous compounds of the seed are mainly in the form of insoluble proteins, such as the glutens of the cereals. Some other proteins, as albumins, are present in small quantities as well as traces of amino-compounds.

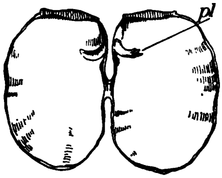

FIG. 144.

336. Germination of Seeds. When a seed is planted in warm, moist soil it first absorbs moisture and swells, then bursts the seed coat, and begins to sprout (Fig. 145). It pushes a shoot upward and a root downward, but until the leaf expands and the root has fairly entered the soil, the young plant derives no nourishment other than water, either from the earth or from the air. It lives on the starch, gluten, mineral matter, and other compounds contained in the seed. The seed, therefore, acts as a storehouse of concentrated food to nourish the plant until it is able to draw its nutrition from external sources. But the substances found in the seed are for the most part insoluble; hence they must undergo a chemical change before they can be taken up into the sap and conveyed along the vessels of the young shoot they are destined to feed. It is so arranged in nature that when the seed first sprouts, there is produced at the base of the germ a small quantity of a white soluble substance called diastase. This substance acts upon the starch, making it soluble in the sap, which is thus enabled to take it up and convey it just as it is wanted, to the shoot or to the root. The starch is thus converted into dextrine and maltose (309). In the oily seeds the mucilage and oil take the place of starch in nourishing the young sprout. The oil is first decomposed into glycerin and fatty acids (303), and these substances are finally converted into carbohydrates.

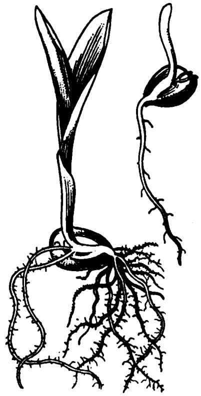
FIG. 145. — Seedlings of corn.

As the sap ascends, the dextrin from the starch is further changed into sugar. This sugar is later changed into cellulose, or woody fiber of the stem and leaf. By the time that the food contained in the seed is exhausted, the plant is able to live by its own exertions, at the expense of air and soil.

In like manner the insoluble protein compounds in the seed are converted into soluble forms similar to the peptones. Some of the soluble proteins are even broken down into amines, which are then in a condition to be transported through the plant tissue and used as building material. These compounds, when they reach the place where they are used, are reconstructed into proteins.

337. Conditions Necessary for Germination. These requirements are the presence of: (1) moisture, (2) oxygen, and (3) heat. If seeds are kept dry, they will not sprout. They will, however, retain their vitality for a long time, and advantage is taken of this fact in the storage of seeds for future use. In the case of agricultural plants, germination is best effected when the soil is moist but not wet.

Likewise seeds which are kept under the proper conditions of moisture and temperature will not germinate if they are not supplied with oxygen. Seeds often fail to grow because they are planted at too great a depth to obtain the needed oxygen from the air. The necessity of oxygen for germination may be shown by a simple experiment. When seeds are put into water those which float are usually the only ones that germinate. Those which sink cannot germinate for lack of oxygen. If a current of air is kept passing through the water all the seeds will germinate. Seeds kept under the proper conditions of moisture and temperature in bottles in which the air is replaced by any inert gas, such as hydrogen or carbon dioxide, will not germinate.

Finally if seeds are supplied with sufficient moisture and air but kept cold, germination will not occur. Seeds sown during cold weather often remain some weeks before sprouting, while those sown in warm weather may germinate in a few days. Seeds of different plants vary in the amount of heat they need for germination. Wheat, oats, and peas will sprout while the ground is quite cool and can be sown early in the spring. Corn requires more heat than oats; while soy beans, cucumbers, melons, and cotton need still more warmth in the soil. Hence the farmer and gardener arrange the time of planting, waiting until the soil has the proper temperature before planting the various kinds of seeds.

338. Carbon Dioxide Exhaled during Germination. The sprouting of the seed starts the hitherto dormant cells into active life and growth, and one result of all life, whether animal or vegetable, is the production of carbon dioxide through respiration (Fig. 146). This accounts for the need of oxygen during germination. The production of carbon dioxide is due to the oxidation of some of the carbon compounds of the seeds.

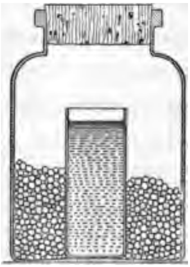

FIG. 146.—**Apparatus showing that germinating seeds produce carbon dioxide.**

339. Roots, Bulbs, and Tubers. Some plants do not bear seeds the first year but go into a resting stage during the winter and produce seeds the following year. Such plants are called biennial. They store the material needed to start growth the second year in enlarged fleshy roots, like the beet; in underground stems or tubers, like the potato; or in enlarged stalks called bulbs, like the onion. The food materials in these storage organs are similar to those found in the seeds; but they are not in such a dry and concentrated form. The chemical changes which these compounds undergo when growth begins in the spring are similar to those which occur during the germination of the seeds.

340. Manufacture of Carbohydrates. If the plantlet produced by the seed is kept in the dark, it remains colorless and grows until the food which was stored in the seed is exhausted. Under normal conditions, however, the leaves of the plantlet become green before the food stored in the seed is completely exhausted, and the green plant has the power of preparing its own food from the simple compounds absorbed from the soil and the atmosphere.

The green coloring material found in the leaves is called chlorophyll, and a microscopic examination of the leaf shows that it is contained in small grains called *chlorophyll bodies* or *chloroplasts*, which are imbedded in the protoplasm of the leaf cell. The chemistry of chlorophyll is little understood, but it is of utmost importance to the plant because its presence enables the protoplasm of the leaf to produce all the organic compounds of the plant. Of special interest is the manner in which the plant manufactures its carbohydrates. The carbon dioxide of the air enters the leaves through the tiny openings or stomata, which are found on the under side, and passes into the air spaces between the cells, and finally into the cells themselves.

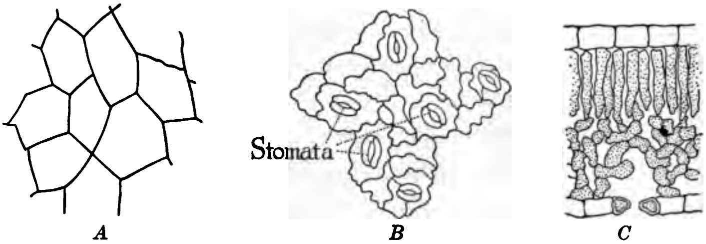

FIG. 147. — **Leaf structure.** *A*, upper surface. *B*, under surface. *C*, cross section.

In the cell the carbon dioxide unites with the water which the plant has absorbed from the soil and forms carbonic acid. Under the action of chlorophyll and daylight the carbonic acid probably breaks up into formaldehyde and oxygen:

$$H_2CO_3 \rightarrow CH_2O + 2\,O.$$

The oxygen is given off through the cell walls into the air spaces and passes out through the stomata into the atmosphere. The formaldehyde is then probably almost instantly changed to glucose, thus:

$$6\,CH_2O \rightarrow C_6H_{12}O_6.$$

The glucose manufactured in this way is then transported to that part of the plant needing new material for growth, where it is changed into cellulose, starch, oil, or other of the numerous plant compounds. But during the daytime the glucose is produced more rapidly than it can be transported, and the cells would be clogged with soluble food if it were not changed to an insoluble form for temporary storage. The plant, therefore, has the power of changing the glucose into starch, thus:

$$C_6H_{12}O_6 \rightarrow C_6H_{10}O_5 + H_2O.$$

At night, when no carbonaceous matter is being formed, the starch is changed back to glucose or other soluble compounds and is transported to other parts of the plant. By morning all the starch has disappeared from the leaves. In testing for starch, the leaf is first boiled to kill the protoplasm and then treated with alcohol to extract the chlorophyll. After this it is placed in a dilute solution of iodine, which gives a violet coloration to any starch which may be present.

341. Daylight Necessary for Carbon Fixation. It has already been stated (105) that daylight is necessary to enable the plant to utilize the carbon of the carbon dioxide. The light waves absorbed by the chlorophyll supply the protoplasm with the energy necessary to enable it to split off the oxygen from the carbonic acid (340). This process of building carbohydrates from water and carbon dioxide is called *photosynthesis;* it is a chemical synthesis (32) by means of light. That light is necessary for the formation of starch can be shown by covering part of a leaf with some opaque substance early in the morning (Fig. 148), and in the afternoon picking the leaf and testing for starch. It will be found that no starch is present in the part that was kept dark, but that starch is abundant in the rest of the leaf. The common garden nasturtium is an excellent plant for this experiment, but any other rapidly growing plant will do. If a few leaves are tested early in the morning they will generally be found to contain no starch, while those tested toward evening will show starch in abundance. Plants kept in the dark will give no test for starch in their leaves, and if allowed to remain in the dark too long they even lose their green color.

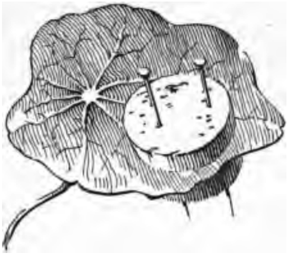
FIG. 148. — **Exclusion of light from part of a leaf.**

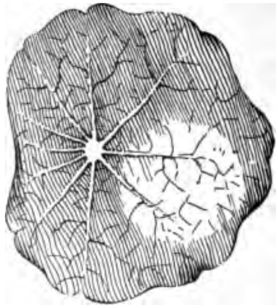
FIG. 149. — **The same leaf tested with iodine, showing absence of starch in part excluded from light.**

The light of the electric arc and, indeed, any white light, can furnish energy for carbon fixation, and it has been suggested that plants might be made to grow more rapidly if supplied with light at night. This has been tried in the greenhouse; but it has been found that while the plants in the houses lighted at night do grow more rapidly, the extra growth does not repay the extra expense of lighting.

At the time of the most color in plants, there is the greatest cell activity and the largest amount of plant tissue is being produced. When a plant ripens, the decline of activity of the cells may be observed by the change in the color of the plant. In corn the lower joints of the stalk turn yellow first, indicating that growth and activity have ceased in those parts. Then the upper leaves become yellow, and finally the husk becomes yellow and inactive. Chlorophyll is one of the principal agents taking an active part in plant growth, and whenever it is destroyed, plant growth is checked.

As the atmosphere contains only 3 parts in 10,000 of carbon dioxide, it may be thought that a larger amount of this gas would enable the plant to make a greater growth. It has been possible, indeed, with plants in small inclosures, to increase the growth by adding carbon dioxide to the air. There is every reason to believe, however, that plants grown in the open never suffer for lack of it and that the size of the crop is always limited by some other factor, and never because of an insufficient supply of carbon dioxide (105–107).

342. Respiration. Every living cell must breathe during the entire period of its active life. Growing plants, therefore, breathe or respire during all of the twenty-four hours. Through respiration the plant takes in oxygen and gives off carbon dioxide. The statement so often seen that "during the night plants take in oxygen and breathe out carbon dioxide, and in the daytime take in carbon dioxide and breathe out oxygen," is not strictly true. The process of photosynthesis should not be confused with respiration. The latter occurs day and night, while the former, which is strictly a manufacturing process and has nothing to do with breathing, takes place only during daylight. It is true that the amount of oxygen given off during photosynthesis is so much greater than that absorbed by respiration that the latter process is obscured during the daytime by the former. The oxygen evolved during six hours of active carbon fixation is as much as would be absorbed by the process of respiration in twenty-four hours. It is evidently true, then, that plants decrease the amount of carbon dioxide in the air during daylight, but that, like animals, they add to it during the night.

343. Changes in the Carbohydrates. The change of the sugar of the plant sap into cellulose, or woody fiber, is more or less observable in all plants. When they are growing fastest the sugar is most abundant, not, however, in those parts that are actually growing, but in those which convey the sap to the growing parts. Thus the sugar of the ascending sap of the maple disappears in the leaf and extremities of the twig, and sugar cane is sweet only a certain distance above the ground, up to where the new growth is proceeding. In the ripening of the ear, the sweet taste so perceptible in young grain gradually diminishes, and finally disappears. The sugar of the sap is here changed into the starch of the grain, which is destined, when the grain sprouts, to be reconverted into sugar for the nourishment of the growing grain.

In the ripening of fruits a different series of changes presents itself. The fruit is at first tasteless, then becomes sour, and at last sweet. In this case, either the acid of the unripe fruit is changed into the sugar of the ripened fruit, or some of the other constituents of the fruit are converted into sugar which disguises the acid.

344. Manufacture of Protein. Carbohydrates are not the only chemical compounds that are produced in the leaf. A very important work of the leaf is the production of proteins, the most complex compounds known. How these compounds are manufactured is not known. Starting with the carbohydrates, and the nitrogen, sulphur, and phosphorus secured from the soil, the leaf cell builds up the very complex protein molecule. Other parts of the plant manufacture protein to some extent; but its production goes on most actively in the leaf, from which it is transported to other parts of the plant.

345. Some Plants Cannot Manufacture Food. Plants not possessing chlorophyll are not able to decompose carbonic acid and produce carbohydrates. Such plants must have their carbonaceous and nitrogenous foods prepared for them. Mushrooms and Indian pipe are examples of such plants. They feed upon the compounds formed by the decaying organic matter in the soil. The fungi that grow on decaying trunks of trees also belong to this class of plants. Some colored plants, like the dodder, are parasites and live on the juices of other plants.

EXERCISES

Ex. 211. Rub 50 kernels of barley in a mortar with a little water. Filter and test the water with Fehling's solution for reducing sugar. What is the result? Put 50 other barley seeds between two thicknesses of moist cotton flannel and place these between two plates as shown in Fig. 150. Stand the plates in a warm place for two or three days or until the sprouts are about half an inch high. Rub the germinated seeds in the mortar as above and test with Fehling's solution. What is the result? What happened to the starch of the seed during germination? To what was this change due? Do similar changes take place in the proteins of the seed?

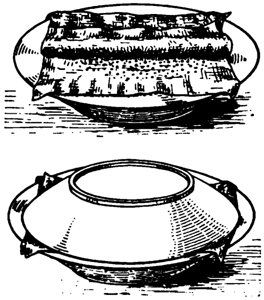

FIG. 150. — A seed tester.

Ex. **212.** What conditions are necessary for the germination of seeds? Boil some water and when it is cold place it in a glass tumbler and drop a few radish or wheat seeds on the surface. Note if there is any difference in the seeds which float and those which sink. Explain.

Ex. **213.** (Teacher) Place an inch of moist sawdust in the bottom of three eight-ounce wide-mouth bottles, and drop a few radish or other seeds on the sawdust. Cork one bottle tightly, and set it aside. Fill another bottle with carbon dioxide by downward displacement and cork it tightly. Fill the third bottle with oxygen by downward displacement and cork it. Watch the bottles for several days and note any difference in germination. Is oxygen necessary for germination? Is heat also necessary? Do seeds vary in the amount of warmth necessary for germination?

Ex. **214.** (Teacher) Place a vial of clear limewater in a wide-mouth bottle and surround the vial with well-soaked seeds (Fig. 146). Cork the bottle tightly and observe the limewater. What change takes place? How do you account for it?

Note. As a check to the above another bottle might be arranged in the same way except that perfectly dry seeds should be used.

Ex. **215.** Pick some leaves from a vigorously growing plant in the afternoon. Boil them in water for a few minutes and then immerse them in hot alcohol to extract the chlorophyll. Dip the leaves in a very dilute solution of iodine. Is there any evidence of starch? How did the starch get there? Explain the production of carbohydrates in the leaf.

Ex. **216.** Cover half of a leaf while on the plant on both sides with black paper or tin foil one afternoon and pick the leaf the next afternoon and test for starch. Is there any difference in the two parts of the leaf? Explain. Pick leaves in the early morning and in the early evening from the same plant. Test for starch in the usual way. Does electric light have the same effect upon plant growth as sunlight? Is there sufficient carbon dioxide in the atmosphere for plant growth?

Ex. **217.** Is it correct to say that " plants breathe out oxygen during the daytime"? Explain what really happens. What is respiration?

Ex. **218.** Are proteins manufactured in the leaf? Can all plants manufacture their own food? How do the fungi obtain their food? How do the parasites obtain theirs?

CHAPTER XXXVII

CHEMISTRY OF PLANT GROWTH (*Continued*)

346. Importance of Water to the Plant. Analysis of the corn plant (334) shows that it contains nearly 80 per cent water. There is also found in the organic matter an amount of hydrogen and oxygen equal to about 10 per cent of the entire plant, and it has been shown that these elements are derived from water (340).

It is evident, then, that the plant obtains about 90 per cent of its substance from water. This statement, however, gives but little idea of the amount of water required by the plant during its period of growth. The leaves of the growing plant are constantly exhaling water. This process is known as transpiration. Very large amounts of water are transpired by plants.

Experiments have shown that while producing one pound of dry matter the plant gives off from 300 to 500 pounds of water by transpiration. A fair crop of corn transpires during the growing season at least 900 tons of water to the acre, or an amount of water that would be equal to a layer that would cover the entire acre about 8 inches deep. The table that appears on the next page gives the average amount of water transpired by some of the common farm crops, as determined at the Wisconsin Experiment Station.

AVERAGE AMOUNT OF WATER USED TO PRODUCE ONE POUND OF DRY MATTER

Crop	Water
Barley	461.1 pounds
Oats	503.9 pounds
Corn	270.9 pounds
Clover	576.6 pounds
Peas	477.2 pounds
Potatoes	385.1 pounds

347. How the Plant Obtains Its Water. All the water used by the plant is absorbed from the soil by the plant roots. The growing ends of the rootlets are clothed with numerous root hairs which are responsible for the absorption of the water needed by the plant. The manner in which the root hairs absorb water from the soil may be illustrated by a simple experiment. Tie a piece of moist animal membrane (hog's bladder will serve the purpose) over the end of a thistle tube (Fig. 151), and when dry cover the edge of the membrane with melted paraffin, to make the joint watertight. Fill the enlarged part of the thistle tube with a strong solution of sugar or salt, and place the tube in a glass of water, sinking it until the level of the liquid in the tube stands at the same height as that in the glass. In a short time


FIG. 151. — Experiment to show osmosis.

the water begins to rise in the tube, and in time flows over the top of the tube. The water passes through the membrane by *osmotic action,* or *osmosis.* Whenever a membrane like this one separates a strong solution from a weak one, there is a decided movement of water from the weaker solution to the stronger, which tends to continue until the liquid is of the same concentration on both sides of the membrane.

Under the microscope the root hairs (Fig. 152) are seen to be long tubelike bladders. These root hairs are filled with cell sap, which is a much more concentrated solution than soil water; hence the water passes into the root hairs by osmotic action. Once in the root hairs the water passes on to the root and stem and leaf, to be utilized in growth or given off by transpiration.

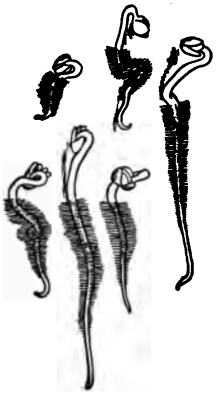

Fig. 152. — **Root hairs of young radish plants.**

348. Functions of Water in the Plant. Water is important to the plant in several different ways. It is first of all the most essential plant food in the sense that it furnishes the material for 90 per cent of the weight of the plant. Water is necessary to dissolve the plant food in the soil and enable it to enter the plant, as will be noted later. It is also necessary for the movement of food within the plant. The food materials absorbed by the roots and those manufactured in the leaves can be transported to the different parts of the plant where they are needed only when in solution in water. Water performs an important function in controlling the temperature of the plant. Chemical processes in the plant cell produce heat, and the excess of heat is removed by transpiration of water through the leaves.

Water is needed also to give stiffness or rigidity to the more succulent parts of the plant. This fact is shown by the drooping or wilting of plants during the hot hours of the day when the water is not furnished by the roots with sufficient rapidity to repair the loss by evaporation from the leaves. In the experiment with the thistle tube (Fig. 151) it was noted that the water was raised to some height in the tube, and consequently the walls of the tube must have been subjected to some internal pressure. The experiment may be performed in another way. Tie a piece of bladder (*A*) over one end of a glass tube (Fig. 153) and fill the tube with a strong sugar solution. Over the other end fasten a piece of thin sheet rubber (*B*), and place the bladder end in a vessel of water. After some time it will be found that the water absorbed has created sufficient pressure to distend the rubber, and if the rubber is punctured with a pin the water will be ejected with some force. The pressure created by osmosis in this way is called *osmotic pressure.* When the plant cells can obtain all the water they need, they are kept distended and rigid by osmotic pressure. When the water is removed faster than it is absorbed, the osmotic pressure is decreased, the cell loses its rigidity and finally the whole plant droops or wilts. The protoplasm does its work properly only when the cell is turgid, and, therefore, wilting is always injurious to the plant.

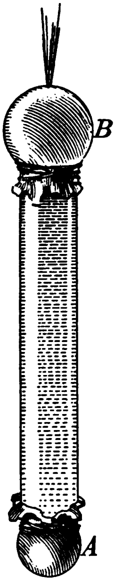

FIG. 153.—**Apparatus to demonstrate osmotic pressure.**

349. How the Plant Obtains Its Mineral Matter. The mineral matter, or ash, of the plant is obtained from the soil. Small quantities of the different mineral substances used by the plant are found dissolved in the soil water. These dissolved materials diffuse into the root hairs by osmosis and then, like the water, pass on to root and stem and leaf to be utilized in plant growth: If none of these substances were used by the plant, this diffusion would continue until there was the same strength of each of the mineral substances in the plant sap and in the soil water. When the plant uses one of these substances, more will come into the root hairs in order to preserve the equilibrium. Thus those substances which are needed by the plant must come in as long as the soil can furnish them in a soluble form.

350. Essential and Non-essential Elements. Although the plant contains nitrogen and the nine ash elements which it obtains from the soil, it does not follow that all of these elements are necessary to its growth. To determine which elements are essential, plants are grown in sand or by the water culture method in such a way that they are supplied with all the elements occurring in plants, with the exception of the one element under investigation. If the plant grows to maturity, the missing element is deemed non-essential; if the plant fails to develop, that particular element is considered to be essential. These experiments indicate that nitrogen, potassium, calcium, magnesium, iron, sulphur, and phosphorus are absolutely essential to plant growth. Toward chlorine, silicon, and sodium plants seem to be indifferent, as they can grow to maturity in the absence of these elements.

Another important fact discovered in these experiments is that one chemical element cannot be substituted for another in plant growth, even when both elements are similar in chemical properties. In the laboratory, for example, sodium and potassium compounds are much alike in their action, and one may be used in place of the other in many reactions;

but sodium cannot take the place of potassium as a plant food (Fig. 154).

351. Roots Dissolve Mineral Substances. The plant not only absorbs what is already soluble in the soil water, but it is capable of making soluble small quantities of the insoluble substances which are present in the soil and which may be needed for plant food. The plant accomplishes this result by means of substances excreted by the roots. If a plant is grown in sand or sawdust placed on a piece of polished marble, it will be found that the prints of the roots are distinctly shown on the surface of the marble (Fig. 155). Pieces of limestone in the soil often show markings due to the solvent action of plant roots.

FIG. 154. — Showing the effect of sodium (*A*) and potassium (*B*) on plant growth.

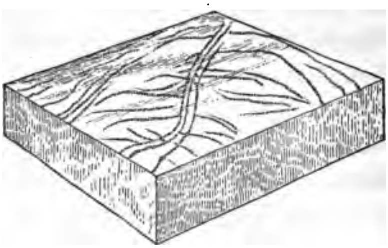

FIG. 155. — Marble corroded by bean roots.

352. The Nitrogen of Plants. This element is largely derived from the nitrates in the soil which enter the root hairs by osmosis. The plants known as the legumes obtain part of their nitrogen from the nodules found on their roots. These are the homes of bacteria that have the power of fixing the nitrogen of the air. They cause the nitrogen to combine with other substances to form compounds which can be utilized as a source of nitrogen for the manufacture of protein. Legumes use the nitrates in the soil when they can obtain them, and only fix atmospheric nitrogen when the supply of nitrate is insufficient.

353. Functions of the Elements in Plant Growth. Carbon, oxygen, and hydrogen are constituents of all the organic compounds manufactured by the plants. These three elements constitute about 98.5 per cent of the mature plant.

Nitrogen is a necessary constituent of the proteins, which play an important part in the formation of protoplasm, chlorophyll, and other compounds. Abundance of nitrogen in the soil is indicated by the bright green color of the leaves.

Potassium is one of the most important elements in plant growth. It is present in greatest amount in the leaves and the actively growing parts of the plant. Apparently it aids in the production of carbohydrates, such as starch and sugar. Abundance of potassium is said to increase the amount of sugar in fruits and in such roots as the sugar beet.

Calcium takes a prominent part in the production of new tissue and in the development of strong cell walls and numerous root hairs. It serves also as a base to precipitate the poisonous oxalic acid which is formed by cell activities.

Magnesium assists in the formation of chlorophyll and the proteins. It is necessary to seed formation, and seeds grown with an insufficient supply of magnesium are often sterile.

Phosphorus is a necessary constituent of some proteins. It is found mainly in the seeds. It increases the yield and hastens the ripening of grain. Many proteins insoluble in water are soluble in the presence of phosphorus compounds.

Iron occurs in the smallest amount of any of the ash elements but is always present in plants. It is necessary for the formation of chlorophyll.

Sulphur is a necessary constituent of most proteins. It is also a part of some of the flavoring oils, such as those found in mustard, onions, cabbage, and horseradish.

EXERCISES

Ex. **219.** Invert a wide-mouth bottle over a potted plant, first covering the soil with waxed paper. (Fig. 156.) What is the source of the moisture that collects? How much water do plants transpire in producing a pound of dry matter? How much water is transpired by an acre of corn?

FIG. 156.—**Experiment showing that water is given off from the leaves of plants.**

Ex. **220.** (Teacher) Perform the experiment described in **347**. How does it illustrate the movement of water into the plant? Germinate some radish seeds between two layers of moist cloth (Fig. 150) and examine the root hairs under the microscope. Are they well designed to absorb water?

Ex **221.** (Teacher) Perform the experiment illustrated in Fig. 153. How does water give rigidity to the plant? Why do plants wilt? Is turgidity of the cell necessary? What are some of the other functions of water in the plant?

Ex. **222.** How does the plant obtain its mineral matter? Describe an experiment to determine which mineral elements are necessary. Name the essential elements.

Ex. **223.** (Teacher) Place a slab of polished marble in a small box and cover with an inch of moist sand. Plant seeds of peas or beans and keep watered. After growth has proceeded for some time wash the marble. Is there proof that the roots dissolved the marble?

Germinate some radish seeds between pieces of blue litmus paper. (Fig. 150.) What effect do the root hairs have on the litmus? Is it probable that roots have power of dissolving mineral food in the soil?

Ex. **224.** What is the source of the nitrogen used by the plant? Discuss the functions of the different elements used in plant growth.

CHAPTER XXXVIII

ENZYMES — DIGESTION — FERMENTATION

354. Enzymes. In studying the germination of seeds (336) it was found that a substance called diastase which has the power of changing starch into maltose is formed in the seed. Diastase belongs to a group of substances known as enzymes. They are the products of living cells but are not themselves living things. Very little is known of the true nature of enzymes or of their chemical action, as none of them have been obtained in a state of absolute purity. They are very complex substances of a protein character. They are soluble in water and glycerin but are insoluble in alcohol. For the purpose of study an enzyme is obtained by pulverizing the tissue, extracting the enzyme with glycerin, and then precipitating it by the addition of alcohol. The material obtained in this way is not a pure enzyme, but contains the enzyme in a concentrated form. Diastase prepared from malt in this way shows in a marked manner the property of transforming starch into maltose.

The enzymes are specific in their action. Diastase, for example, converts starch into maltose, but it has no effect on other substances; and in like manner each enzyme acts on one particular substance, producing in it a definite change. The enzymes behave like catalytic agents in that they themselves undergo no permanent change, but under proper conditions can cause an almost indefinite amount of chemical change in the substance upon which they act. Thus one part of diastase can change at least 2000 parts of starch into maltose without any of the enzyme itself being destroyed.

355. Malt is a good example of the commercial utilization of enzymic action. Malt is produced from barley by soaking the grain in water for some time and then spreading it in thick layers upon the floor of a warm room. Germination takes place, and when the sprouts are about one half inch long, the grain is heated sufficiently to kill the embryo, and then dried. The sprouts are removed and sold as a cattle feed under the name of *malt sprouts*. The remaining grain is known as malt. The germinating process makes the diastase active, and if the malt is now placed in warm water, the starch of the grain is converted into maltose, which may be changed by the action of yeast into alcohol, as is done in the manufacture of beer, whisky, and ordinary alcohol. Since the amount of diastase in the barley is capable of changing a large amount of starch into maltose, other starchy materials, such as corn and rice, are frequently added to the " mash." The residual grain, after all the starch has been made soluble and thus removed, is dried and sold as cattle feed under the name of *dried brewers' grains*, or *dried distillers' grains*, according to whether it comes from the brewery or the distillery.

356. Digestion. The process by which the insoluble food materials are made soluble so they can be absorbed into the blood of animals is called digestion. It is in large part, if not wholly, the result of the action of several enzymes. Digestion takes place in various parts of the alimentary canal, notably in the mouth, stomach, and small intestines.

357. Digestion in the Mouth. The food is first ground into fine particles by mastication so that the digestive juices can act upon it to better advantage. During this process the food is thoroughly mixed with the saliva, which contains an enzyme known as *ptyalin*. This enzyme is much like diastase in its action and changes the starch of the food into maltose.

$$\underset{\text{starch}}{2\,C_6H_{10}O_5} + H_2O \rightarrow \underset{\text{maltose}}{C_{12}H_{22}O_{11}}.$$

The normal saliva is slightly alkaline, and ptyalin can act only in an alkaline solution. No constituents of the food other than starch are acted upon in the mouth, and not all the starch is rendered soluble. The food material, thoroughly moistened, passes into the stomach, where the next change takes place.

358. Digestion in the Stomach. In man, the horse, and the pig there is but one stomach, but in the ruminants, like cattle and sheep, there are four stomachs (Fig. 157) or rather four compartments to the stomach. Animals of the latter class chew the cud. The food is passed from the mouth into the first and second compartments of the stomach, and is then forced back into the mouth for further mastication; then it is swallowed again and passed through the third stomach into the fourth for final digestion. The storage in the first and second stomach and the repeated mastication of the food merely serve to grind the food completely and to prepare it thoroughly for digestion.

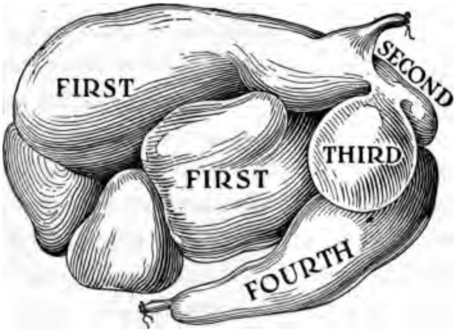

FIG. 157. — The four main divisions of a ruminant's stomach.

In this way these animals are able to digest fibrous material to a much greater extent than other animals, such as the horse. The true gastric digestion in the case of ruminants takes place in the fourth stomach.

The glands in the wall of the stomach secrete a digestive fluid, called gastric juice, which, unlike the saliva, is acid in reaction and contains about 0.2 per cent of hydrochloric acid. It also contains two enzymes — pepsin and rennin. The pepsin acts upon the insoluble proteins and gradually converts them into peptones, which are soluble and diffusible. Pepsin acts only in an acid solution. Rennin, the other enzyme of the gastric juice, acts on the casein of milk, causing it to coagulate or curdle, a process the necessity of which is not understood. The coagulated casein is then dissolved by the pepsin. Rennin is especially abundant in the stomach of the young, and the commercial rennet used in cheese making is prepared from the stomachs of young calves. No food constituents save the proteins are acted upon by the stomach enzymes, and they are not completely digested but in part pass on into the small intestine.

359. Digestion in the Intestine. When the food reaches the small intestine it comes in contact with the intestinal and pancreatic juices. These fluids have an alkaline reaction and contain several enzymes.

Trypsin is an enzyme that acts upon the proteins which escape digestion in the stomach. It is more energetic in its action than is pepsin. It acts only in an alkaline solution.

Amylopsin is a pancreatic enzyme that acts on starch, converting it into maltose, and is more energetic in its action than ptyalin.

Steapsin, or lipase, is an enzyme that acts upon the fats of the food. It hydrolyzes fats into glycerin and fatty acids

(298), which is probably the first step in their digestion. Steapsin and amylopsin are active only in alkaline solution.

360. Bile is a fluid secreted by the liver and discharged into the small intestine together with the pancreatic juice. It is a thin liquid, with a bitter taste, and is very alkaline. It varies in color from greenish-yellow to reddish-brown, the shade depending on the animal. No enzymes have been discovered in the bile, but its presence decidedly increases the power of the pancreatic enzymes.

The various food materials that have been changed into soluble compounds by the action of the digestive enzymes are absorbed from the small intestines and ultimately find their way into the blood to be transported to the part of the body where they are needed. The part of the food that is not changed into soluble compounds passes on into the large intestine and is finally excreted in the feces, which, therefore, represent in a general way the undigested food.

361. Other Enzymes. The enzymes mentioned in this chapter are only a few of the many whose existence is known, and new ones are being constantly added to the list. The pineapple is known to contain an enzyme that digests protein. Some enzymes bring about oxidation by causing the union of substances with the oxygen of the air. Such enzymes are called *oxidases*. The brown coloration which appears on the cut surface of an apple or other fruit is said to be caused by the action of an oxidase. The enzymes studied herein assist in breaking complex substances into simple bodies, but there are undoubtedly enzymes that produce opposite results. The synthesis of starch in the leaves and the production of proteins, as well as many other processes of the plant and animal body are thought by some investigators to be dependent upon the presence of enzymes.

362. Fermentation is a term applied to changes in organic substances that are brought about through the growth of microscopic plants, such as yeast, molds, or bacteria. The production of alcohol by the action of yeast is the best-known example of fermentation. The souring of milk, the change of cider to vinegar, and the decay of organic substances are examples of fermentation caused by bacteria. It was formerly thought that the growing cells themselves produced the chemical changes incident to the fermentation, but many investigators now believe that the yeasts, or bacterial cells, produce enzymes that are really responsible for the chemical changes in the fermenting material. It was found that a sample of yeast, for instance, that was ground in such a way as to rupture every cell and thus destroy its life, still had the power of producing alcoholic fermentation. It is now said that yeast contains at least two enzymes: namely, *invertase*, which has the power of inverting cane sugar, and *zymase*, which converts invert sugar or maltose into alcohol and carbon dioxide.

EXERCISES

Ex. **225.** Crush 20 malted barley grains in a mortar. Transfer to a test tube, add 15 cc. of water, and allow the mixture to stand twenty-four hours. Filter off the solution and add to a bottle containing 100 cc. of starch solution made as follows: rub one gram of starch with 10 cc. of water until smooth and then pour on 100 cc. of boiling water and allow the liquid to cool. Allow the mixture to stand another twenty-four hours. Test a portion of it for starch. State the result. Test another portion with Fehling's solution. State the result. What change has taken place in the starch? What caused this change? What is the active principle of the malt? What are enzymes? Are they specific in their action? Explain. Are they catalytic agents? What commercial use has the enzymic action of malt? What are dried brewers' grains?

Ex. **226**. Fill a test tube one third full of your saliva. If the saliva does not flow freely chew a piece of paraffin. Add to the saliva an equal volume of starch solution prepared as in the last exercise. Place the test tube in a cup of water at blood heat for an hour. Test a portion of mixture for starch. If the starch has not all disappeared, allow the mixture to stand another hour. What change has the saliva caused in the starch? How did you test for starch? What is the enzyme of the saliva? How is the food affected in the mouth? Is the saliva acid, or alkaline, or neutral?

Ex. **227**. What enzymes are found in the stomach? Review Exercise 200. Repeat that experiment, substituting a tablespoonful of ground lean meat for the white of the egg. State the result. Does pepsin act best in an acid or alkaline solution? Compare with ptyalin. What action does the rennin of the gastric juice have on milk?

Ex. **228**. What three enzymes are found in the intestines? State the action of each. What effect does the bile have on digestion?

Ex. **229**. Explain what is meant by fermentation. Give examples. Do enzymes play any part in fermentations? Are enzymes concerned in any processes other than digestion and fermentations? What causes the brown coloration of the cut surface of an apple? (*Note.* If an apple is cut in two and placed in a bottle containing sulphur dioxide for a short time, the cut surface will not turn brown when exposed to the air as the sulphur dioxide destroys the enzyme oxidase.) Are enzymes ever supposed to take part in building up complex compounds?

CHAPTER XXXIX

PRINCIPLES OF NUTRITION

363. Uses of Food. The animal body uses the foods for the following purposes: (1) to repair the waste of the system; (2) to supply heat; (3) to furnish motion; (4) to provide the materials needed for the increase of flesh by growth or fattening; (5) to make special products, such as milk, eggs, feathers, wool, and hair. The animal body may in many ways be compared to the gasoline engine or other "prime motor." The gasoline engine requires two things for its operation: (1) sufficient repair material to keep its working parts in running order, and (2) a supply of fuel in proportion to the work to be done. The same two things are needed by the animal; namely, repair material and fuel.

364. Repair Material. The repair material for any machine must be of the same kind as that of which the machine is made. Protein is the characteristic ingredient of the animal mechanism; for the muscles with which the animal does its work are largely composed of protein, and this material is broken down and destroyed at a fairly uniform rate by the operation of the animal machine. Since the bodily machinery is running all the time, whether any external work is being done or not, this loss is going on continually. The body differs from the engine in being self-repairing, but as the animal does not have the power to manufacture proteins, it is absolutely dependent for its repair material on the proteins of its food. This protein is needed for two purposes.

First, it is necessary for repair material in the strict sense; namely, to make good the wear and tear of the body machinery. The amount needed for the purpose is small, and is not materially greater when the animal is doing work than when it is not. A second purpose for which protein and ash are needed in the growing animal is to furnish the material for enlarging its body. Protein is necessary also to enable the animal to manufacture milk, eggs, hair, wool, and other special products; for all of these contain proteins, which the animal must obtain from its food.

365. Food as a Source of Repair Material. The value of a food or feeding stuff as a source of protein evidently depends in the first place on the amount of protein which it contains. Beans, containing 23 per cent of protein, are, other things being equal, a better source of protein than corn, which contains only 10 per cent. Since the protein of the food must be capable of being digested by the animal, the most valuable source of repair material is the food or feeding stuff containing the largest amount of digestible protein (368).

366. Fuel or Energy Materials. The animal requires heat to maintain the body temperature and energy to do its work. The source of this heat and energy is the food which the animal digests, and which is oxidized in its body. Since the animal machinery is running continually, it requires a constant supply of fuel material, the amount necessary depending upon the amount of work done. This consists chiefly of the carbohydrates and fats of the food, although if more protein is fed than is required for repair and construction purposes, it may be used as fuel. The unnecessary use of protein as fuel material is wasteful, as protein is ordinarily much more expensive than are carbohydrates and fats.

367. Fuel Value, or Energy Value, of Foods. The different foods and food constituents are not all of equal value as sources of energy. It will be found convenient to have a means of comparing the different foods and feeds, and the best basis for such a comparison is the relative energy values of these materials. Anything which has the capacity to do work is said to possess energy. The fuel of the engine and the food of the animal possess energy, since they enable the engine or the body to do work. This energy is stored up as latent energy, and when the fuel is burned in the engine, or the food is oxidized in the body, this latent energy is set free and part of it is converted into the work, the rest escaping as heat. The value of a fuel depends on the amount of this latent energy it contains, and this can be determined by burning the substance to convert the latent energy into heat, and then measuring the heat produced.

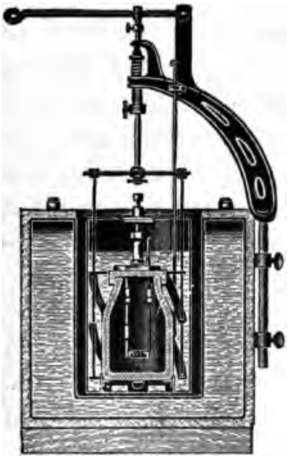
FIG. 158. — A calorimeter.

The fuel value of a food is determined by burning a weighed quantity of the food in a calorimeter (Fig. 158). This is a metal vessel or bomb, which is immersed in a vessel of water. The heat produced by the burning substance warms the water, and the rise in temperature is determined by a thermometer. Various units have been employed in measuring heat. Perhaps the oldest and most common is the calorie (spelled with a small c) which is the amount of heat necessary to raise the temperature of one gram of water one degree centigrade. This unit is so small that it has long been customary in discussing foods to use the large Calorie (spelled with a capital C), which is equivalent to 1000 small calories, or, in other words, is the amount of heat required to raise the temperature of one kilogram (1000 grams) of water one degree. In discussing the feeding of farm animals even the large Calorie is found to be an inconveniently small unit, and some writers make use of the *therm*, which is equivalent to 1000 large Calories; that is, it is the amount of heat required to raise the temperature of 1000 kilograms of water one degree centigrade.

A pound of either carbohydrates or protein when burned produces 1860 large Calories of heat, or 1.86 therms. A pound of fat, which is a much more concentrated fuel, produces 4225 Calories, or 4.23 therms of heat. It will be seen that the pound of fat produces approximately two and one fourth times as much heat as an equal weight of carbohydrates or protein. The following examples serve to illustrate the great variation in the energy value of different foods:

FOOD	ENERGY VALUE PER POUND
	Calories
Cabbage	170
Potatoes	380
Eggs	720
Wheat flour	1660
Cheese	1990
Butter	3600

It has been shown that only the protein that is digested is of any use to the animal body, and it is equally true that only the digested part of the food can supply the animal with energy. The undigested part, which passes off in the excrement, represents that part of the energy that cannot be utilized by the animal machine. To determine how much energy a food will furnish, it is necessary to know what proportion of the food is actually digested; and that can be determined only by experiments with animals.

368. Digestion Experiments. The digestibility of a food is determined by a carefully conducted digestion experiment. The animal, or man, is fed for a few days on the food to be investigated, the food having been analyzed to determine its content of protein, fat, carbohydrates, fiber, and other constituents; and the amount eaten is carefully recorded. The feces, which represent the undigested part of the food, are collected, weighed, and analyzed. The difference between the undigested nutrients and the total amounts in the food consumed is the amount digested, which is then calculated on a percentage basis. For example, suppose the analysis of the food shows that the animal consumed four pounds of protein during the experimental period, and that the feces contained one pound of protein. The animal, then, was able to digest three pounds of protein out of the four pounds consumed. In other words, three fourths, or 75 per cent, of the protein was digestible. In the same way the percentage of digestibility may be worked out for the fats, ash, and each of the other nutrients of the food. Some authors use the expression *coefficient of digestibility* to designate the percentage of a nutrient which can be digested. In the above-assumed example the coefficient of digestibility for protein is 75.

369. Available Energy of Foods. Only that part of the food which is digested can furnish energy to the animal machine, and not all of this energy can be utilized by the body. Some of the digested portion of the food fails to undergo complete oxidation in the body and is excreted in the liquid excrement. The energy that can be utilized is called the available energy of the food. It is determined by finding the total number of Calories in the food, and subtracting therefrom the caloric value of the feces as well as the caloric value of the compounds found in the liquid excrement.

370. Net Energy of Foods. In the process of digestion, particularly of coarse fodders, a part of the energy of the food is used to separate the real fuel material from the relatively large proportion of useless material in the food. The energy thus used up in carrying on the process of digestion is not available for other purposes. It is possible to determine the approximate amount of energy required by the animal in order to chew the food and digest it, and this amount subtracted from the available energy gives the net energy of the food. In the case of coarse fodders a large part of the available energy is used in digestion, leaving comparatively little net energy. Less energy is required to digest grains, and thus a larger proportion of their energy can be used for other purposes by the body. The total fuel value of one pound of timothy hay, for example, is 1751 Calories, but its net energy value is only 335 Calories. The total fuel value of corn meal is 1709 Calories and its net energy value is 888 Calories. Therefore, while the total fuel value of these two substances is not very different, there is a marked contrast in the amount of net energy which they furnish.

If the fuel materials supplied in the food are just adequate to the work to be done, they are all burned up as a source of power. If more are supplied than are immediately needed, the body is able to store the surplus for future use. Most of the surplus fuel is converted into fat, which, therefore, is the reserve fuel of the body. In fattening, the body is accumulating a surplus against future needs. If the food later becomes insufficient, this store is drawn upon and the animal becomes thin. Similarly, in growth and milk production, the animal sets aside a part of the supply of both repair and fuel material in its food for its own growth or for the use of its young. Man takes advantage of these tendencies of the animal to store fat and meat, and to produce milk, and diverts the resulting products to his own use as repair and fuel material for his own body.

EXERCISES

Ex. **230.** What five general uses does the animal make of its food? Why is the animal body likened to a prime motor engine? What furnishes the repair material for the animal body? What other use does the animal have for protein? Does the value of food for repair depend on the total protein or upon the digestible protein?

Ex. **231.** What are the chief energy supplying substances in the food? How is the total energy value of a food determined? What is the measure of heat? What is meant by the small calorie; the large Calorie; the therm? How much heat is produced by burning a pound of protein or carbohydrate? A pound of fat? Do foods vary greatly in their energy value?

Ex. **232.** Does digestibility affect the amount of energy which the food will supply to the animal? How is a digestion experiment conducted? What is meant by the available energy of the food? How is it determined?

Ex. **233.** What is meant by net energy? In what way does it differ from available energy? Can all the available energy be utilized to do work? Explain. Can all the net energy be used to do work? Is some energy required to prepare the food and to digest it? How does the proportion of net energy from seeds compare with that from fodders? If the net energy supplied is more than is used, what becomes of the excess? Can the animal use its own body fat to supply energy upon occasion? Is energy needed to produce milk and growth?

CHAPTER XL

FEEDING FARM ANIMALS

371. Balanced Rations. It has been noted that food supplies the animal with repair material and with energy, both of which it needs to carry on its various functions. The amount of repair material and energy required by the animal depends upon the following factors: whether the animal is growing, or is working, or is producing milk. A ration that will supply the animal with protein and energy in just the proportion in which it needs them is called *a balanced ration.* A knowledge of the food requirements of animals and of the method of calculating balanced rations should be of value to the practical feeder.

372. The maintenance requirement of the animal is used as the basis for calculating the balanced ration. Since the animal machine cannot be stopped when it is not in active use, it requires a continual supply of food. The amount of food that is required simply to support the animal is designated as the maintenance requirement. It is the amount required simply to maintain the animal when it is doing no work and producing nothing. It represents the least amount on which life can be maintained. A large animal needs more food for maintenance than does a small one, although the difference is not exactly proportional to the weight, but appears rather to be approximately proportional to the body surface of the animal. The proper maintenance requirements have been determined by experiment for different kinds of animals, of various ages and weights. The following table gives the figures for cattle of four different weights.

MAINTENANCE REQUIREMENTS FOR CATTLE

WEIGHT	DIGESTIBLE PROTEIN	NET ENERGY
Pounds	Pounds	Therms
750	0.40	4.95
850	0.45	5.60
1000	0.50	6.00
1250	0.60	7.00

373. Requirements for Growth. The amount and nature of the food consumed should vary with the period of growth as well as with the size of the animal. Rations for young growing animals should contain proportionately more digestible protein and less energy value than rations for mature animals. This is because more food is required for building purposes in the early stages of growth than in the later stages, when the demand is more for heat and energy. When an excess of fats and starchy foods is given to young animals, there is a tendency to produce poor muscular tissue and premature fattening. The ash of the food is also very important to young animals, for it is during the growing period that the bones are built up. A table showing the protein and energy requirement for growing animals will be found at the end of this chapter.

374. Requirements for Work. The performance of work by the animal calls for an additional supply of energy in the feed. Animals when doing medium or heavy work also require more protein than do those at light work.

REQUIREMENTS FOR THE WORKING HORSE OF 1000 POUNDS

CHARACTER OF WORK	DIGESTIBLE PROTEIN	NET ENERGY
	Pounds	Therms
For light work	1.0	9.80
For medium work	1.4	12.40
For heavy work	2.0	16.00

375. Requirements for Fattening. When the animal consumes more energy-making foods than it can utilize, it stores the surplus energy as fat. To fatten animals, then, they are fed abundant rations high in net energy value. It is estimated that about 3.5 therms in addition to the maintenance requirement are needed by the animal for each pound of gain in weight during the fattening period.

376. Requirements for Milk Production. Of all forms of animal production that of milk is the most variable and most influenced in its amount by the feed supply. Milk is the natural food of the young, and, as it is the only food of the very young animal, it contains the protein, the ash, and the energy necessary to its growth. When a cow is producing milk, she must have in addition to her maintenance ration an amount of food sufficient to enable her to put protein and energy materials into her milk. To produce a pound of average milk requires 0.05 pound of digestible protein and 0.3 therm of energy, which must be added to the maintenance ration for each pound of milk the cow produces.

377. Dry Matter in Rations. It has been found by experiment that it is necessary for cattle to have a certain bulk in their feed. They do not thrive so well if the feed is too concentrated, but on the other hand there might be such a thing as having a ration which is too bulky. The best indication of bulk in the feed is the dry matter which it contains. In a general way it may be said that an animal weighing 1000 pounds should be given from 20 to 30 pounds of dry matter a day, the exact amount not being very important if kept within these limits. On the farm, where hay and fodder are abundant, it is usually easy to obtain a ration that is sufficiently bulky.

378. The Ash of the Ration. The ash, or mineral matter, of the feed is important; for the animal could not live very long if there were no mineral matter whatever in the feed. It is especially important to young, growing animals, as they are building up bones which are composed very largely of mineral matter. As the animal grows older, it needs less ash in the feeds, since the bones are no longer growing in size.

379. Calculating a Balanced Ration. With the data of this chapter at hand it is possible to calculate rations suited to the various needs of the domestic animals. To illustrate the method of calculation, it is assumed that a ration is needed for a cow weighing 850 pounds and producing 20 pounds of average milk each day. The feeds are to be selected from the table given at the end of this chapter.

By referring to page 332 it is seen that a cow weighing 850 pounds requires for maintenance 0.45 pound of digestible protein and 5.60 therms of energy. For the production of 20 pounds of milk of average quality there would be required according to the figures given in paragraph 376:

Digestible protein (0.05 × 20)	1 pound
Net energy value (0.3 × 20)	6 therms

The total feed requirements for a day for such a cow are, therefore, 1.45 pounds of digestible protein and 11.60 therms of net energy.

The problem is to find a mixture of feeds that will give these amounts of protein and energy. As the coarse feeds grown on the farm are usually the cheapest, they should be used as far as possible. First, corn silage and clover hay may be tried for *roughage,* as the coarse feeds are called, and corn meal and wheat bran for the more concentrated feeds. It is necessary to start with the best guess possible as to the amounts of each feed to use and make a table showing the results as below:

Ration	Dry Matter	Digestible Protein	Net Energy Value
Pounds	Pounds	Pounds	Therms
Corn silage 22	5.63	0.26	3.64
Clover hay 6	5.08	.32	2.08
Corn meal 5	4.46	.34	4.44
Wheat bran 2	1.77	.20	.96
Total	16.94	1.12	11.12

A comparison of these totals with the requirement of 1.45 pounds of protein and 11.60 therms of energy shows that the ration is slightly low in energy and considerably so in protein. If the addition of some feed high in digestible protein, say 1½ pounds of gluten feed, is made the ration stands thus:

Ration	Dry Matter	Digestible Protein	Net Energy Value
	Pounds	Pounds	Therms
In feeds named above . . .	16.94	1.12	11.12
In 1½ pounds gluten feed . .	1.38	.30	1.19
Total	18.32	1.42	12.31

This ration gives more nearly the correct amount of digestible protein, but has a surplus of energy, which would probably tend to fatten the cow instead of increasing the flow of milk. The energy in the ration should be reduced without decreasing the amount of protein. If one pound of corn meal, which supplies chiefly energy, is omitted, and replaced by one half pound of gluten feed, the ration is as follows:

RATION	DRY MATTER	DIGESTIBLE PROTEIN	NET ENERGY VALUE
Pounds	Pounds	Pounds	Therms
Corn silage 22	5.63	0.26	3.64
Clover hay 6	5.08	0.32	2.08
Corn meal 4	3.56	0.27	3.55
Wheat bran 2	1.77	0.20	0.96
Gluten feed 2	1.84	0.40	1.59
Total	17.88	1.45	11.82

The ration now agrees very closely with the computed requirements. This example will serve to illustrate the method of calculating all rations; for the same method will apply to the rations for fattening cattle, for horses, for sheep and swine, and for chickens, if the standards for each kind of animal are known. By proceeding in the manner described, with a little patience a ration corresponding as closely as is necessary to the standard requirements can be calculated. Experience makes it possible to guess pretty closely the first time, and the computation soon becomes easy.

380. Individuality. The standard requirements, of course, are for average animals, but it is well known that some animals require more feed than the average and some less. The wise feeder, therefore, uses the standards with this fact in mind, and in addition to calculating his standard ration makes an individual study of each cow in his herd, feeding her any amounts for which she will give profitable returns. Even such a feeder, however, needs the standard requirements as a starting point in his study.

381. Palatability of Feeds. There is another factor in feeding animals which is quite as important as that of balancing the ration; namely, the matter of the palatability of the ration. In order to give the best results the food should be relished by the animal. The experienced feeder strives to compound a ration that carries the proper proportion of protein and energy and is pleasing to the animal's taste.

382. Older Feeding Standards. The feeding standards that have been most commonly used in the past by writers on the feeding of animals are those known as the Wolff-Lehman standards. These standards, instead of being based on the daily requirements of the animal for digestible protein and energy, are based on the theory that the animal must have a given weight of dry matter each day, together with a definite amount of the three digestible nutrients — protein, carbohydrates, and fat. The following table gives a few of the feeding standards according to the Wolff-Lehman tables:

DAILY REQUIREMENT OF DIGESTIBLE NUTRIENTS FOR EACH 1000 POUNDS LIVE WEIGHT OF ANIMAL

ANIMAL	DRY MATTER	PROTEIN	CARBOHYDRATES	FAT
	Pounds	Pounds	Pounds	Pounds
Cows giving 22 pounds of milk daily	29	2.5	13.0	0.5
Fattening cattle	30	2.5	15.0	0.5
Sheep	23	1.5	12.0	0.3
Horses, medium work	24	2.0	11.0	0.6
Fattening swine	36	4.5	25.0	0.7

The method of calculating a ration according to these standards is exactly the same as the one described on page 335, except that in the case of the Wolff-Lehman standards there are four factors to be balanced, while in the other case only two items, protein and energy, are considered. In the case of the Wolff-Lehman standards, the best possible guess is made as to the feeds that will fit the standard, and then others are added or subtracted from the ration, as described in the foregoing example, until a mixture is obtained that agrees very closely with the amounts of dry matter, protein, carbohydrates, and fat as stated in the standard. These older standards are being replaced by those based on protein and energy. Tables giving all the Wolff-Lehman standards, as well as the percentages of the different digestible nutrients in the common feeding stuffs, may be found in the larger works on the feeding of farm animals.

383. The nutritive ratio of a feed or a ration is the proportion between the digestible protein and the sum of the digestible carbohydrates and fat contained therein. To find the nutritive ratio the fat is multiplied by 2.25, because it has 2.25 times the food value of the carbohydrates, and the result is added to the carbohydrates. The sum is divided by the digestible protein, the quotient being the nutritive ratio. Thus, the standard for a horse at medium work calls for 2.0 pounds digestible protein, 11.0 pounds carbohydrates, and 0.6 pound of fat.

$.6 \times 2.25 = 1.350; \quad 1.350 + 11.0 = 12.35; \quad 12.35 \div 2.0 = 6.17$

The nutritive ratio, therefore, is 1 to 6.17.

Recent investigations indicate that the ratio between protein and the other nutrients is not so important as it was first thought to be, provided that the animal is supplied with sufficient repair material or protein.

TABLES: ESTIMATED REQUIREMENTS PER DAY AND HEAD FOR GROWING ANIMALS

CATTLE				SHEEP			
Age Months	Live Wt. Pounds	Digestible Protein Pounds	Net Energy Therms	Age Months	Live Wt. Pounds	Digestible Protein Pounds	Net Energy Therms
6	425	1.30	6.0	6	70	0.30	1.30
12	650	1.65	7.0	12	110	0.23	1.40
18	850	1.70	7.5	18	145	0.22	1.60
24	1000	1.75	8.0				

CONSTITUENTS IN 100 POUNDS OF FEEDING STUFFS

Feeding Stuffs	Dry Matter Pounds	Digestible Protein or Repair Material Pounds	Net Energy Therms
Coarse Feeds			
Corn silage	25.6	1.21	16.56
Alfalfa hay	91.6	6.93	34.41
Clover hay	84.7	5.41	34.74
Corn fodder	57.8	2.13	30.53
Corn stover	59.5	1.80	26.53
Oat hay	84.0	2.59	36.97
Timothy hay	86.8	2.05	33.56
Oat straw	90.8	1.09	21.21
Mangels	9.1	.14	4.62
Grains			
Barley	89.1	8.37	80.75
Corn	89.1	6.79	88.84
Corn and cob meal	84.9	4.53	72.05
Oats	89.0	8.36	66.27
Rye	88.4	8.12	81.72
Wheat	89.5	8.90	82.63
By-products			
Dried brewers' grains	92.0	19.04	60.01
Cottonseed meal	91.8	35.15	84.20
Distillers' grains	93.0	21.93	79.23
Gluten feed	91.9	19.95	79.32
Gluten meal	90.5	33.09	78.49
Linseed meal	90.8	27.54	78.92
Malt sprouts	89.8	12.36	46.33
Dried sugar-beet pulp	93.6	6.80	60.10
Wheat bran	85.1	10.21	48.23
Wheat middlings	84.0	12.79	77.65

EXERCISES

Ex. **234.** What is meant by a balanced ration? By the maintenance requirement of an animal? Do large animals have a greater maintenance requirement than small ones? Do growing animals need proportionately more protein in their foods? Why is the ash of the food very important to young animals? When animals are doing work, do they need more energy producing materials than when not working? Do they also need more protein? What kind of foods do fattening animals need — those high in protein or in energy?

Ex. **235.** How much protein and energy are needed to produce a pound of milk? Is the bulk of the food of any moment? What can you say about the effect of palatability on the value of the ration? How important is it to study the individuality of the animal in feeding?

Ex. **236.** How is an animal's daily ration calculated?

Ex. **237.** Calculate a ration for a 1000-pound dairy cow that gives 30 pounds of milk a day. Use corn fodder and alfalfa hay for roughage and any of the foods in the table on page 339 for concentrates. Calculate the protein and the energy in any ration used on your home farm and note whether it agrees with the feeding standards given in this chapter.

Ex. **238.** Calculate the ration for a fattening ox weighing 1000 pounds according to the Wolff-Lehman tables on page 337, using corn stover and clover hay for roughage. Which method of calculating is the simpler? For analyses of feeding stuffs see Farmers' Bulletin, No. 22, U. S. Department of Agriculture, or any of the larger texts on feeding animals.

CHAPTER XLI

HUMAN FOODS

384. Food Requirements of Human Beings. The principles of human nutrition are exactly the same as those for domestic animals. Men and women need protein for the repair of tissue and for energy to enable the body to do its work, and these two necessities are furnished by the food, just as in the case of the lower animals. The food requirements of human beings are indicated by dietary standards that have been worked out by investigators. The following standards of daily requirements have been prepared by Atwater:

Character of Work	Protein	Fat	Carbohydrates	Calories
	Pounds	Pounds	Pounds	
Man with little exercise . .	0.20	0.20	0.66	2450
Man with light work . . .	0.22	0.22	0.77	2800
Man with moderate work .	0.28	0.28	0.99	3520
Man with hard work . . .	0.39	0.55	1.43	5700

A woman is supposed to require eight tenths of the protein and energy needed by a man; and children require an amount about proportional to their size and weight. In general terms it may be stated that according to this standard a man of average size and doing average work requires about one fourth of a pound each of protein and fat, one pound of carbohydrates, and 3200 calories of energy daily. Some writers think this standard too high and hence likely to result in overfeeding.

385. Calculating a Balanced Ration. The nutritive value of human foods varies just as it does in the case of the feeds for the farm animals. The method of calculating a ration that conforms to the standard is exactly the same as that used in calculating the ration for cattle or horses (382). The following combination given by Snyder serves as an example of a day's ration which would meet the general standard for a man doing average work.

FOODS	AMOUNT OF FOODS PER DAY	PROTEIN	FAT	CARBOHYDRATES	CALORIES
	Ounces	Pounds	Pounds	Pounds	
Ham	4	0.04	0.09	—	480
Eggs (2)	—	0.03	0.02	—	136
Bread	8	0.05	0.01	0.28	650
Butter	2	—	0.11	—	450
Potatoes	12	0.02	—	0.14	285
Milk	16	0.04	0.04	0.05	325
Sugar	2	—	—	0.12	200
Beef stew . . .	4	0.04	0.05	—	250
Oatmeal	2	0.02	0.01	0.09	230
Corn meal . . .	4	0.02	0.01	0.18	420
Totals . . .		0.26	0.34	0.86	3426

This ration contains somewhat less carbohydrates and more fat than the standard and furnishes a little less energy, but it is close enough for practical purposes.

386. Practical Use of Dietary Standards. It is neither practicable nor necessary to undertake to prepare the meals each day in exact conformity to a dietary standard. An occasional study of the foods served in the family to ascertain how closely they conform to the standard is desirable, because such a study gives a basis for modifying the diet, if necessary to make it supply the proper amount of protein and energy. It is practicable, after a study of the composition of the different foods, to make combinations that will provide in a general way the right proportions of protein and energy, and to avoid combinations that are too high in protein or that carry an excess of energy. In other words, it is practicable to avoid a combination that includes several foods high in protein, or one made up of several foods high in energy value and low in protein. In combining foods to form balanced rations it is well to remember that lean meats, fish, dried beans and peas, oatmeal, and nuts are substances high in protein. Fat pork products and other fat meats, cheese, butter, oils, and lard supply fats in large proportions. Potatoes, rice, corn meal, cereals, sugars, cornstarch, and tapioca are high in carbohydrates. Wheat flour and the other foods prepared from wheat are moderately high in both protein and carbohydrates and low in fat.

The more recent studies of foods indicate that the subject is much more complicated than was formerly supposed. It is not sufficient merely to balance the food so as to provide the required amount of protein, fat, carbohydrates, and energy. The proteins differ among themselves in character; hence, the ration should contain a variety of proteins.

It has been discovered also that most common foods contain very small quantities of a water-soluble substance which is essential for the maintenance of normal body conditions. This substance has not been isolated and its nature is not understood. It is lacking in certain prepared foodstuffs such as polished rice, commercial starch, pure sugar, and fats.

Another of the important discoveries of modern chemistry is the fact that there is a fat-soluble substance, called vitamine, which seems to be absolutely necessary to growth (Fig. 159). The nature of this substance is unknown, but it is known that it is not so widely distributed as is the water-soluble substance mentioned above. It is found in milk, especially in the milk fat, in egg yolk, in some meats, and in the green leaves of plants. It is not found in seeds except in the germ, nor in oils such as olive and cottonseed oils. A lack of this fat-soluble substance in the food prevents the growth of young animals even when abundantly supplied with food. The fact that it is present in milk is one of the reasons why milk is so valuable a part of the diet of growing children.

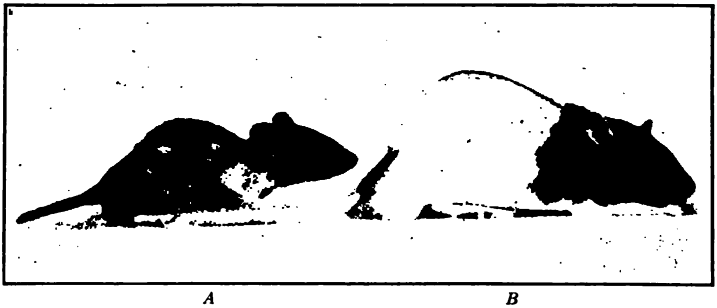

Fig. 159. — (*A*) **Rat without and** (*B*) **rat with fat-soluble vitamines.**

There is a difference in the character of the mineral elements, or ash constituents, of foods. In some foods the acid-forming elements are in excess, while in others the basic or alkaline elements predominate. Recent investigations indicate that it is desirable that the diet should contain a slight excess of basic elements. The foods that furnish alkaline mineral substances in relatively large proportions are tubers, leafy vegetables, fruits, and milk. The cereals, meat, fish, and eggs contain the acid mineral elements in excess. It would seem, therefore, that such a combination as bread and milk is a logical one, since in this combination the alkalinity of the ash materials of the milk overcomes the acidity of the cereal ash.

Palatability is even more important in human foods than in animal feeding stuffs. Food should be so prepared as to appeal to the individual by its appearance and flavor, since pleasure in eating undoubtedly plays a part in insuring a regular and normal sequence of digestive process.

387. Digestibility of Foods. The term digestibility has been used in two different ways by physiologists and chemists: (1) to designate the completeness of the process of the digestion of the food; and (2) to designate the fact that the food is digested without causing distress or discomfort during the process. Some confusion has arisen from this double meaning. Cheese, for example, is very completely digested; but since it is commonly considered to be hard to digest, that is, to cause distress after eating, it is often said to be indigestible. Bread, which is digestible according to the second use of the word, is not so completely digested as cheese.

Some of the factors affecting the digestibility of foods are the following: (1) *Individuality of the person.* Some people can easily digest foods that cause great discomfort to others. (2) *Mechanical condition of food.* When the food is in good mechanical condition it is more easily acted upon by the digestive juices. (3) *The combination of foods.* The way in which foods are combined is of importance, as some foods seem to aid in the digestion of others. (4) *Method of preparation.* The method of preparing or cooking foods exerts an influence on their digestibility. Cooking changes both the physical and the chemical condition of the food, and influences the ease of digestion more than the completeness of the process. The chief advantages of cooking are probably the development of a more pleasing flavor, and, more important still, the complete sterilization of food that may contain injurious bacteria or other organisms.

388. Food Adjuncts. Spices, such as allspice, nutmeg, cinnamon, and ginger, are not foods, but may or may not be useful adjuncts to the diet. They stimulate the appetite and promote secretion of the gastric juice, which may be beneficial or may induce overeating. Such substances are called condiments. Vinegar is a condiment and is used largely because of the pleasant acidity of its acetic acid.

Coffee or tea have little or no nutritive value but are used because of the stimulating effect of the alkaloid caffein which they contain. Cocoa contains a similar alkaloid but has some nutritive value as well, because of its fat content.

389. Fresh Fruits. Considered strictly from the chemical standpoint, fruits seem to have little food value, as they are low in both protein and energy. Fresh fruits contain from 85 to 95 per cent water, a fraction of 1 per cent of fat and protein, and only 5 to 10 per cent of carbohydrates. They contain, however, acids, organic salts, and other substances which are believed to have a beneficial effect on the system, and they doubtless often stimulate the appetite for other food. Fruits also add to the attractiveness of the diet, and the appearance of the food is a matter of considerable importance. The ash of fruits is rich in potassium, calcium, and iron salts, all of which are valuable to the body. Dried fruits, such as dates, raisins, and figs, are foods in the more restricted sense, as they furnish a large amount of digestible nutrients. Dates, especially, form a large part of the diet of certain Oriental people.

390. Dietary Studies. A dietary study takes into consideration the cost and amount of nutrients consumed by individuals or families. It is an investigation in which men and women and human foods are used instead of farm animals and animal feeds. In a dietary study, the amounts of nutrients in the edible portion of the foods are determined by chemical analysis or calculated from the tables of composition. These studies take into consideration the cost of the material as well as the amounts of protein and energy used by each person or group. Such studies show that lack of knowledge in regard to the values of foods has frequently resulted in whole families being underfed, not from necessity, but from lack of judgment in the selection of foods. It too often happens that choice of foods is made wholly on the basis of palatability, instead of on the cost of the nutrients and the kind of work to be performed. Dietary studies show that for long periods the best results are obtained when the foods are combined in such a way as to furnish the different nutrients in approximately the amounts given in the dietary standards. By studying the diet it is often possible to reduce the cost of the food without impairing its nutritive value if indeed the nutritive value is not actually increased.

391. Food Fads. Much of the matter that has been written on the subject of foods is wholly without scientific basis, even when it pretends to be scientifically presented. Vegetarianism, or the exclusive use of vegetable foods, is advocated by some people; but there is no scientific evidence that mankind is benefited by an exclusively vegetable diet. The long-continued experience of the human family on a mixed diet of cooked meats and vegetables is evidence that such a diet is healthful, and there are many indications that the best diet is one that contains a reasonable amount of animal protein. Many so-called vegetarians obtain this animal protein through the use of milk and eggs; but of course in such conditions the diet is no longer strictly vegetarian. It is probably true that American families use more animal protein than is necessary, a practice which, while it may not be injurious, results in an unnecessarily expensive diet.

The exclusive use of raw foods is another food fad that has no scientific basis, and it is not surprising that the cult has a comparatively small following. Unfortunately a large part of the literature on foods has been written by dyspeptics or by people who have prepared foods of some kind to sell.

392. What to Eat. Probably the only advice that can safely be given is about as follows: (1) make the diet agree approximately with the ratios between protein, fat, and carbohydrates as given in the dietary standard; (2) provide a part of the protein from animal sources; (3) consume moderate amounts of such a balanced food; (4) avoid all foods that personal experience has shown to produce discomfort; (5) have as much variety as possible in the diet, including the use of green and succulent vegetables; (6) combine coarse or laxative foods with those that are more completely digested; (7) use fresh fruits abundantly, if possible; (8) use condiments or the stimulating beverages such as tea and coffee only moderately; (9) remember that the highest priced foods are often the least nutritious, and that there is no close relation between cost and food value; (10) use only foods that have been carefully protected from bacterial contamination.

Millions of dollars have been spent by manufacturers in the last few years in advertising the many brands of breakfast foods on the market. In general, it may be said that these breakfast cereals have no greater nutritive value than the grains from which they were prepared. Breakfast cereals made from corn are equal to the same weight of corn meal, for example, and those made from wheat are no more valuable than the wheat itself. The price paid for each pound of actual nutrients in the breakfast foods is several times the cost of the same nutrients in corn meal or wheat flour.

EXERCISES

Ex. **239.** What is meant by a dietary standard? What is the approximate dietary requirement of a man doing average work? How should you calculate a balanced ration for a man? What practical daily use can be made of dietary standards? Name some foods high in protein; some high in fats; some high in carbohydrates.

Ex. **240.** Make a list of the foods used on your home table for each meal for a day. Does the combination appear to be too high in protein? In energy materials? Would a little change improve it? For the composition of the common foods, see Farmers' Bulletin 142, or Office of Experiment Station Bulletin 28, U. S. Department of Agriculture.

Ex. **241.** In what two ways is the term digestibility used? What are some of the factors which affect digestibility of foods? Tell what you can about the food value of fruits.

Ex. **242.** Of what value is a dietary study? How is it conducted? What do these studies sometimes show? Study some of the bulletins of the United States Department of Agriculture on dietary studies.

Ex. **243.** State ten practical points to be observed in deciding what to eat. What can you say about food fads? What can you say about the value of the prepared breakfast foods?

CHAPTER XLII

MILK AND ITS PRODUCTS

393. Secretion of Milk. Milk is a fluid secreted by the mammary glands of all animals that suckle their young. It contains in a palatable and easily digested form all the nutrients necessary for the nourishment of the young animal. Market milk in this country is almost entirely cow's milk, but the goat and the water buffalo are important sources of milk supply in some other countries. In the wild state the cow produced only sufficient milk to nourish the calf until it could subsist upon other food, but under domestication the secretion of milk by the cow has been greatly increased by careful selection and liberal feeding.

394. Composition of Milk. The essential constituents of milk are water, fat, protein, sugar, and ash or mineral salts. The average composition of cow's milk is shown in the following table:

Average Composition of Cow's Milk in Per Cents

Cow's milk	**100**	Water	87.2				
		Solids	12.8	Fat . . .	3.75		
				Protein .	3.2	Casein . .	2.5
						Albumin .	0.7
				Sugar . .	5.15		
				Ash . . .	0.7		

There is considerable variation in the composition of the milk from different cows. The most variable constituent of the milk is its fat content. Some cows produce milk with as little as 2 per cent of fat, while other cows have been known to produce milk containing as much as 8 or 9 per cent of fat. The other constituents of the milk are fairly constant in amount even in milks that vary greatly in fat content. The quantity and the quality of the milk produced by a cow depend upon a number of factors. Certain breeds, such as the Jersey and Guernsey, as a rule produce a relatively small quantity of milk, which is high in fat; while the Holstein and Ayrshire breeds give larger quantities of milk, which is low in the percentage of fat. Individual animals within any of these breeds differ in the quantity and the quality of the milk they produce. The kind and the amount of feeds the animal receives have an influence on the amount of milk produced, but apparently have no effect upon the quality of the milk. The richness of a cow's milk seems to be natural to her and is not affected by the feeds she eats, although the quantity of milk she produces may be so affected. A cow usually produces the most milk per day within a month after the calf is born, and the amount gradually decreases until the secretion ceases as the cow goes dry. As the amount of milk decreases, the percentage of fat is slightly increased. The first milk drawn from the udder at any milking is much poorer in quality than the last. The first often tests as low as 1 per cent of fat and the last as high as 8 or 9 per cent fat.

395. Fat of Milk. This fat occurs in the form of small globules (Fig. 160) which can be seen only under the microscope. The globules average about one six-thousandth of an inch in diameter. The size of the globule averages larger in the milk of Jerseys and Guernseys than in that of Ayrshires or Holsteins. Any sample of milk, however, contains globules that vary greatly in size. The fat globules are held in suspension by the other solids of the milk in the form of an emulsion. The fat of milk is commonly called butter fat. It differs chemically from other fats in that it contains about 5 per cent of butyrin, the glycerin salt of butyric acid, $H \cdot C_4H_7O_2$ (299). The fat has a much higher commercial value than any other part of the milk, and consequently the price of milk is usually based on its fat content.

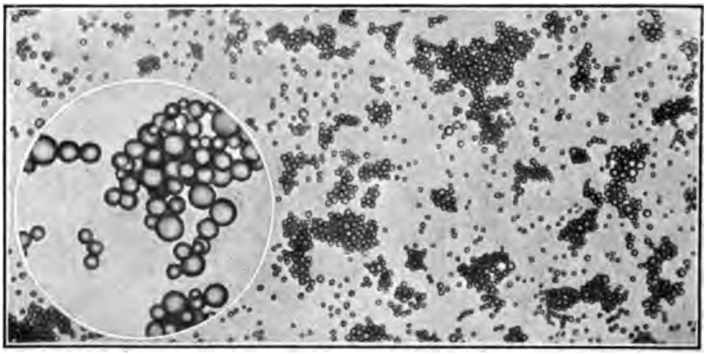

FIG. 160. — **Appearance of milk under the microscope, showing groups of fat globules. In the circle the fat globules are more highly magnified.**

396. Casein. The principal protein of milk is casein and it gives skim milk its bluish-white color. When acid is added to milk the casein separates in the form of a curd. The natural curdling of milk is caused by the lactic acid formed in the milk when it sours (290). The casein can also be separated from the milk by means of rennet, which is a preparation made from the stomachs of very young calves and contains the enzyme rennin (358). When this material is added to milk, the casein separates in the form of a sweet curd, which firmly incloses nearly all the fat that the milk contains. This curd forms the starting point in the manufacture of cheese. Junket tablets contain dried rennet. In making junket just enough rennet is used to coagulate the casein and change the milk into a jellylike mass without causing the casein to separate.

397. Albumin of Milk. This is much like the white of egg. Like all albumins it is soluble in water, and as it is not precipitated by acid it remains in solution when the casein is separated either by acid or rennet. Boiling the clear liquid, or whey, which remains after the casein is removed coagulates the albumin and causes it to separate in white flakes. The tough scum which forms on the surface of milk when it is boiled is composed largely of coagulated albumin.

398. Sugar of Milk. Lactose or milk sugar has the formula $C_{12}H_{22}O_{11}$ (310). This sugar occurs only in milk. In commerce it is found as a fine white powder with a mild, sweet taste. It is about one sixth as sweet as cane sugar. It is readily acted upon by the bacteria in the milk and is changed into lactic acid:

$$C_{12}H_{22}O_{11} + H_2O \rightarrow 4\ \underset{\text{lactic acid.}}{H \cdot C_3H_5O_3}$$

It is this acid that makes milk sour and causes the casein to curdle or separate. When about 0.4 per cent of lactic acid is present the milk acquires a sour taste, and when the amount reaches 0.6 to 0.7 per cent it begins to curdle. Ordinarily the acid will not develop beyond 0.9 per cent.

399. Ash of Milk. The mineral matter left after burning off the organic matter is the ash of milk. It contains all the compounds necessary to build the bony structure of the growing animal. The most important elements in the ash are calcium, phosphorus, iron, potassium, magnesium, and sulphur. The mineral matter is probably largely combined with the casein and albumin in the milk.

400. Milk and Bacteria. Since milk is a complete food and is in a liquid form, it is an ideal medium for the growth of bacteria, molds, and other organisms. Bacteria, especially those which cause the souring of milk, are so widely distributed that it is difficult to keep them out of milk, and for this reason the production of milk which is pure enough for human consumption requires more care than any other work on the farm. Careful attention should be given to the surroundings, to the cleanliness of the cow, to the actual process of milking, and to the utensils used, so as to prevent as far as possible the introduction of bacteria into the milk. Disease-producing germs should be especially guarded against, as many of them grow rapidly in milk. Diseases have often been spread in this way.

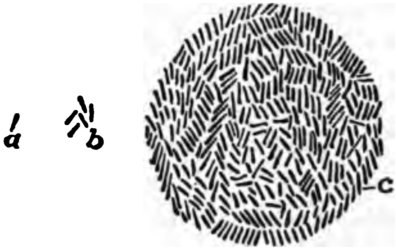

FIG. 161. — **Showing the rapid development of bacteria in milk.** *a*, a single bacterium; *b*, increase in 24 hours when properly cooled; *c*, increase when not cooled.

It is practically impossible to keep all acid-forming bacteria out of the milk; but if the milk is cooled immediately upon being taken from the cow, their growth will be retarded and the milk will keep sweet longer than when this precaution is not taken (Fig. 161).

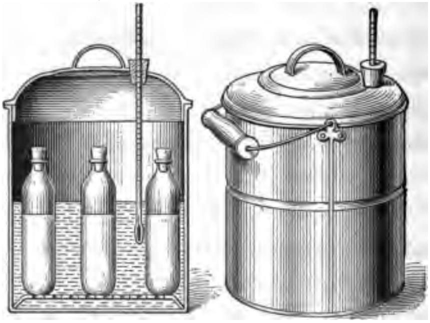

FIG. 162. — **Pasteurizing apparatus.**

Milk that has been boiled to kill the bacteria is said to be *sterilized.* Such milk if carefully stored will keep a comparatively long time but has a cooked taste. Sometimes the milk is heated to 145° F. for from 20 to 40 minutes and is then rapidly cooled, when it is said to be *pasteurized* (Fig. 162). Such milk does not have the taste of boiled milk and if stored in a cool place will keep sweet for a few days. This process kills most of the active bacteria and thus delays the souring.

Antiseptic materials, such as boric acid, salicylic acid, and formalin are sometimes added to milk to preserve it; but this practice is illegal. Any antiseptic that will prevent the growth of bacteria is unfit for use in any food intended for human consumption. Such materials are especially harmful in milk to be used for infant feeding.

401. Cream and Skim Milk. The fat of milk is lighter than the liquid portion of the milk. The specific gravity of the fat is about 0.9, while the rest of the milk has a specific gravity of about 1.036. The fat globules being lighter tend to rise to the surface, where they form a layer known as cream. The larger the globule the more rapidly it rises; so the milk of the Jersey and Guernsey breeds creams more easily than that from the breeds with smaller fat globules, such as the Ayrshire and Holstein. The smaller the fat globule, the larger is its surface in proportion to its volume, and, consequently, the greater the resistance to its rise. Cream may contain from 12 to 50 per cent of fat. That part of the milk that remains when the cream is removed is known as skim milk. It differs from whole milk in containing only 0.1 to 0.4 per cent of fat. When milk is allowed to stand in deep or shallow pans until the cream collects on the surface (24 to 36 hours), the cream is said to have been separated by gravity. Milk held in deep cans which are allowed to stand in cold water creams more completely than when set in shallow pans.

402. Cream Separators. At the present time most of the cream is produced by skimming the milk by the centrifugal cream separator. In this machine centrifugal force generated by a rapidly revolving bowl takes the place of gravity and acts with a much greater force. As the milk flows into the revolving bowl in a continuous stream, it is acted upon by centrifugal force and flies to the outer wall of the bowl. The skim milk being heavier than the cream is forced out and against the side of the bowl, forcing the cream toward the center. By providing suitable outlets the skim milk can be directed into one channel and the cream into another. The skimming is more complete if the milk is first warmed to about 85° F. If separated as soon as it is milked the temperature is right without artificial heating.

The skimming by the separator is much more complete than by any of the gravity methods. A properly working separator will leave not to exceed 0.1 per cent of fat in the skim milk, while gravity creaming may leave from 0.4 to 0.8 per cent. The average per cent composition of separator skim milk is about as follows:

Water	90.54
Fat	0.10
Sugar	4.94
Proteins	3.53
Ash	0.89

Skim milk contains all the materials found in whole milk with the exception of fat. It has a high food value, and its use as a human food deserves more consideration than it has received in this country. Its value as a food for farm animals has long been recognized.

The use of the separator has the added advantage that the skimming is done while the milk is fresh and sweet and that therefore both the cream and the milk are in the best condition for use as foods. Separators vary in capacity from the small hand machines that will handle 150 pounds of milk an hour to power machines with a capacity of 4000 pounds or more an hour.

403. Butter. When cream is agitated for some time at the right temperature, the fat globules unite into larger and larger aggregates, and the fat finally separates in irregular masses of butter. The process of agitating the cream is known as churning. Many different kinds of churns are in use, the types varying from the old-fashioned dasher churn to the more modern barrel-shaped churns.

FIG. 163. — Churning butter in Palestine.

The proper temperature for churning ranges from 50 to 58 degrees F., varying with different samples of cream. Churning should cease when the granules of butter are the size of wheat grains. The butter is then washed with cold water to remove the buttermilk, and is finally worked to remove excess of water. During the working salt is added gradually, although a few American and many European markets demand unsalted butter. American butter contains on the average from 83 to 85 per cent of fat, 12 to 15 per cent of water, and from 2 to 4 per cent of salt.

404. Ripening of Cream. Butter made from perfectly sweet cream is considered by many to have an insipid taste. In order to develop the flavor preferred by the general market and to facilitate the churning, the cream is allowed to become slightly acid before it is churned. This is known as the ripening of the cream. The cream is properly ripened when it contains about one half per cent of lactic acid. Since many of the bacteria that find their way into milk produce undesirable flavors in the cream, in order to insure the proper fermentation it is customary to add a pure culture of the desired bacteria, known as a starter. In many creameries the cream is first pasteurized to kill most of the native bacteria, and the starter is then added. This method makes it much easier to control the flavor of the butter.

405. Buttermilk is the liquid remaining in the churn after the separation of the butter from the cream. It contains about 90 per cent water, 3.5 per cent proteins, 0.5 per cent fat, 4.0 per cent sugar, and 0.5 per cent lactic acid. The finely divided condition of its proteins makes it readily digested. The mildly acid taste of buttermilk is pleasing to some people and very distasteful to others. Buttermilk is growing in popularity as a food and a beverage to such an extent that many factories are now producing large quantities of artificial buttermilk. This is made by adding a starter to skim milk and allowing the proper degree of acidity to develop; when this acidity is properly developed, the milk is churned to break the curd into fine particles such as exist in natural buttermilk. If, as is sometimes done, a little cream is added, the product is known as creamed buttermilk.

406. Condensed Milk and Milk Powders. Condensed milk is prepared by evaporating the water from milk in a vacuum pan until the milk is reduced to about one third or one fourth of the original volume. The evaporation is carried on under reduced pressure so that the milk need not be heated to a sufficiently high temperature to impart to it a cooked flavor. The condensed milk is then sealed in cans and sterilized by exposing the can to superheated steam. In some brands of condensed milk the sterilization is omitted, and about 40 per cent of cane sugar is added to prevent fermentation. The unsweetened condensed milk is quite commonly called *evaporated milk* or *evaporated cream* to distinguish it from the sweetened condensed milk.

Milk powders are prepared by evaporating the milk to complete dryness. A good sample of milk powder is as fine as flour, and when stirred up with water makes a mixture having the properties of milk. Most of the milk powders are made from skimmed or partially skimmed milk. It is said that milk powder prepared from whole milk will not keep on account of the large amount of fat that it contains.

407. Cheese. A great number of varieties of cheese are found in the market, most of which are prepared by the action of rennin (358) on milk. The variety most largely used in this country belongs to the type known as Cheddar cheese and is commonly known as American cheese.

In making American Cheddar cheese, after the milk has been allowed to develop about 0.25 per cent of lactic acid a small quantity of rennet extract is added, and the milk is kept at a temperature of about 85 degrees F. In about thirty minutes the milk sets into a firm, jellylike curd. This curd is then cut into small cubes with specially devised knives, and the temperature of the vat is raised to 100° F. At this temperature the curd shrinks quite rapidly and more acid develops and is absorbed by the curd. After one or two hours the whey is drawn from the vat and the curd mats into a solid mass. After some time it is passed through a mill to shred it and is then salted and pressed into molds. The cheese is then placed in the curing room for a period in order to ripen and develop flavor.

When the cheese is first made, it is tasteless and very tough and rubbery, and is not readily digested. After a period of ripening the cheese becomes soft and plastic and develops a flavor which increases in intensity with the age of the cheese. Practically none of the nitrogenous compounds of the new cheese are soluble in water, while in old cheese over half of these compounds are soluble. In other words the cheese has been partially digested during the ripening process. A good American cheese contains about 26 per cent of proteins and 33 per cent of fat.

Very slight differences in the amount of rennet added, in the temperature at which the milk is set, in the amount of acid developed, or in the temperature and moisture of the curing room produce marked differences in the appearance and flavor of the cheese; consequently there are almost endless varieties of cheeses on the market, varying from the very soft Brie and Camembert to the firm Cheddar and Swiss cheeses. In some varieties, such as Roquefort, molds are added to produce a special flavor. Goat's milk and sheep's milk are also sometimes used in cheese making.

The whey left in cheese making contains most of the albumin and the sugar originally present in the milk. The albumin is sometimes coagulated by heat and used to make an albumin cheese. In some factories the sugar is recovered by evaporation and is sold under the name of lactose, or milk sugar, for use in medicine.

Cottage cheese, or pot cheese, is made from the curd of sour milk without the use of rennet. The curd is firmed by heat and the whey drained off through a cloth strainer. Salt is added and the product is improved by the addition of cream, and sometimes by the use of nutmeg, caraway, or other spices. A similar cheese is made from buttermilk.

408. Other Milk Products. In some creameries the skim milk is utilized for the production of dried casein. Sulphuric acid is used to curdle the milk and the casein is pressed and then dried. The casein thus prepared is used as a sizing for paper and in the manufacture of certain cements. It is used also in massage creams and in some cosmetics. Casein so treated as to make it very hard is used in making billiard balls and knife handles under the name of artificial ivory.

Koumiss is a beverage made from milk. Yeast and sugar are added to the milk, and the ensuing fermentation results in the production of a small percentage of alcohol. Koumiss was originally made in Russia from mare's milk.

Bulgaris milk is usually made from skim milk by the action of a species of bacteria, known as *Bacillus bulgaricus.* These bacteria cause the milk to become sour without the separation of the casein. Bulgaris milk tastes very much like buttermilk, but often has a higher acidity.

EXERCISES

Ex. **244.** What are the constituents of milk? What constituent is most variable in the amount present? How does the breed affect the amount of fat in the milk? Does the feed change the percentage of fat in a cow's milk? Can the quantity of milk be changed by varying the feed? Which contains the most fat — that first drawn from the udder or the strippings?

Ex. **245.** Examine a drop of milk under the microscope. In what form is the fat present in the milk? How large are these globules? What effect does the breed of cow have on the size of the globule? What is the characteristic compound of butter fat? Why does the percentage of fat usually determine the value of the milk?

Ex. **246.** To some skim milk add a little vinegar or acetic acid. What is the curd which forms? To another portion of the milk add a little rennet extract or a little pepsin. What happens? How much casein does ordinary milk contain?

Ex. **247.** Pour off the clear liquid from the milk to which the vinegar was added in the last exercise and heat nearly to boiling. Does albumin separate? How does albumin compare with the white of egg?

Ex. **248.** (Teacher) To a quart of skim milk add just enough acid to coagulate the casein. Heat the whey to coagulate the albumin. Skim off the albumin and evaporate the liquid to dryness to obtain milk sugar. What are the properties of milk sugar? What change takes place when the milk sours? How much lactic acid must develop in the milk before it curdles?

Ex. **249.** What mineral elements are found in milk? Of what use are they to the young animal? To obtain the ash from milk place the milk in an evaporating dish, add sufficient nitric acid to coagulate the casein, evaporate to dryness, and burn off the organic matter.

Ex. **250.** Why do bacteria grow so readily in milk? How do the bacteria get into the milk? What precautions are necessary to produce clean milk? How does immediate cooling of the milk affect its keeping quality? What is meant by sterilized milk? By pasteurized milk? Why do they keep longer than ordinary milk? Why should all antiseptic materials be avoided in milk? Test a sample of milk for formaldehyde as follows: place a little milk in a test tube; incline the tube and pour a little sulphuric acid such as is used in the Babcock test down the side of the tube so that it will underlie the milk. If formaldehyde is present, a purple ring will appear at the junction of the milk and acid.

Ex. **251.** When milk stands why does the cream form at the top? How much fat does cream contain? What is skim milk? What is meant by gravity creaming? Explain the principle of the cream separator. Does the separator remove more fat from the milk than the gravity process of creaming? How much fat is left in separator skim milk? What can you say of the food value of skim milk?

Ex. **252.** How is butter made? How much fat does butter contain? What is meant by the ripening of cream? Why is cream ripened? What is buttermilk? To what is its acid taste due? How is artificial buttermilk made?

Ex. **253.** How is condensed milk made? Examine a can of condensed or evaporated milk. Dilute with twice its bulk of water. Does it taste like ordinary milk? Why is sugar sometimes added to milk? How are milk powders prepared? Why are they usually made from partly skimmed milk? If possible obtain a sample for inspection.

Ex. **254.** Outline the method of making American Cheddar cheese. Rub a piece of well-ripened Cheddar cheese as large as a walnut with a test tube full of water and filter. To the filtrate add a little tannic acid solution. Are there any soluble proteins in the cheese? How were they formed?

Ex. **255.** How many different kinds of cheese are on the local market? Why can so many kinds of cheese be made from one kind of milk? How is cottage cheese made? Is any rennet used in this case? What other uses are made of casein? What is koumiss? Bulgaris milk?

CHAPTER XLIII

TESTING MILK

409. Need of a Test. It has been stated that the fat is by far the most valuable constituent of the milk. Butter fat, as it is more commonly termed, often sells for over 30 cents a pound while the rest of the milk may be purchased for a fraction of a cent a pound. Since milk varies in fat content from as low as 2 per cent to as high as 8 per cent, it is unfair to pay the same price per hundredweight for all kinds of milk. The creameries and factories now almost universally pay for milk on the basis of its fat content, and consequently a quick, easy, and accurate method of determining the fat in milk is very important. Such a test is also of great value to the farmer in enabling him to determine which cows in his herd are being kept at a profit to him. A test recently made of a dairy herd showed that one cow was yielding a profit of over $80.00 a year, while another was being kept at a loss of $15.00. By testing all the cows and replacing those found to be unprofitable with cows showing a better test the profits of a dairy farm may be increased materially.

410. The Babcock Test. The most practical method for testing milk for fat is the one invented by Dr. Stephen Moulton Babcock in 1890 and known as the Babcock test. This test has the advantage of being simple and easily manipulated, and long use has demonstrated its great accuracy. The principle of the test is the separation of the fat by centrifugal force in such a way that it can be measured.

Eighteen grams of milk are treated with sulphuric acid, which dissolves the casein of the milk and thus facilitates the separation of the fat. The mixture is then whirled in a centrifugal machine, and the fat collects on the surface of the liquid. To make the test the apparatus shown in Figs. 164 and 165 is necessary.

The test bottle is so made that the graduations on the neck give the percentage of fat when the sample of milk weighs 18 grams. The small divisions represent 0.2 per cent. The pipette is graduated at 17.6 cubic centimeters, because experience has shown that when filled to the mark the pipette will deliver just 18 grams of average milk, and it is much easier to measure the milk than to weigh it.

411. Sampling the Milk. Milk creams so readily that the greatest care is necessary to get a fair sample of the lot to be tested. The entire amount should be thoroughly mixed by pouring several times slowly from one vessel to another, or by stirring the milk and at the same time taking care that no air is introduced into it. Many errors are made by not securing a correct average sample. When the milk is thoroughly mixed, the small end of the pipette should be

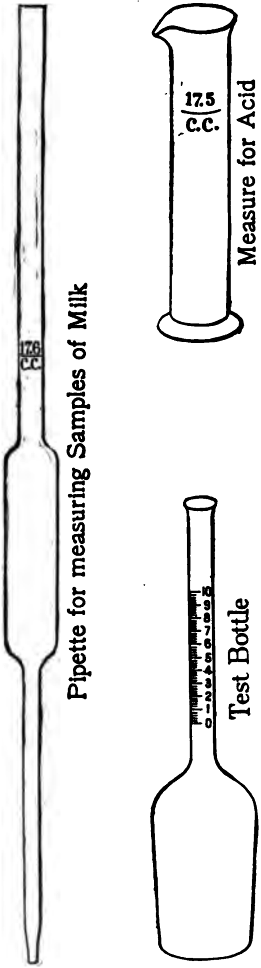

Fig. 164.

placed in it and the tester should suck on the upper end until the milk stands in the tube about an inch above the mark. Then he should quickly slip a finger over the upper end of the pipette, easing the finger slightly until the level of the milk is exactly at the mark. Next he should place the lower end of the pipette in the neck of the bottle (Fig. 166), which is held in an inclined position, and allow the milk to run into the bottle. Finally, when the pipette has completely drained, he should blow out the last drop of milk.

FIG. 165. — Centrifugal machine.

412. Adding the Acid. The acid used should be what is known as commercial sulphuric acid and should have a specific gravity of about 1.81 to 1.83. The pure acid used in the chemical laboratory is too strong and if used will char the fat. The acid measure should be filled to the mark (17.5 cc.) and the acid should be poured carefully into the test bottle, which is held in an inclined position. If this is done properly, the acid runs down the side of the bottle and stands in a distinct layer beneath the milk. When all the bottles are in readiness, the milk and the acid are mixed by gently rotating the bottle until the curd is all dissolved and the liquid is uniform in color.

413. Whirling the Bottles. When the milk and acid are thoroughly mixed the bottles are placed in the centrifuge (Fig. 165) and whirled for five minutes. Directions are furnished with most machines which indicate the number of turns of the handle a minute to produce the proper speed for the bottles.

With a wheel 20 inches in diameter the correct speed is 900 revolutions per minute. If the diameter is greater or less than 20 inches, the speed must be varied so as to produce centrifugal force equivalent to 900 revolutions of the 20-inch wheel. At the end of the five minutes the bottle is filled with hot water to the base of the neck and whirled for two minutes. Then it is filled with hot water until the top of the fat column is near the top of the graduation on the neck of the bottle, and whirled for one minute. Experience has shown that the three whirlings and two additions of water are necessary to insure a clear column of fat which will permit of accurate reading. Unless there are sufficient bottles to fill the machine, care must be taken to balance the weight exactly by arranging the bottles evenly on opposite sides of the machine. If the number of tests made is odd, an extra bottle filled with water is used to preserve the balance. Usually two tests are made of each milk, one to check the other.

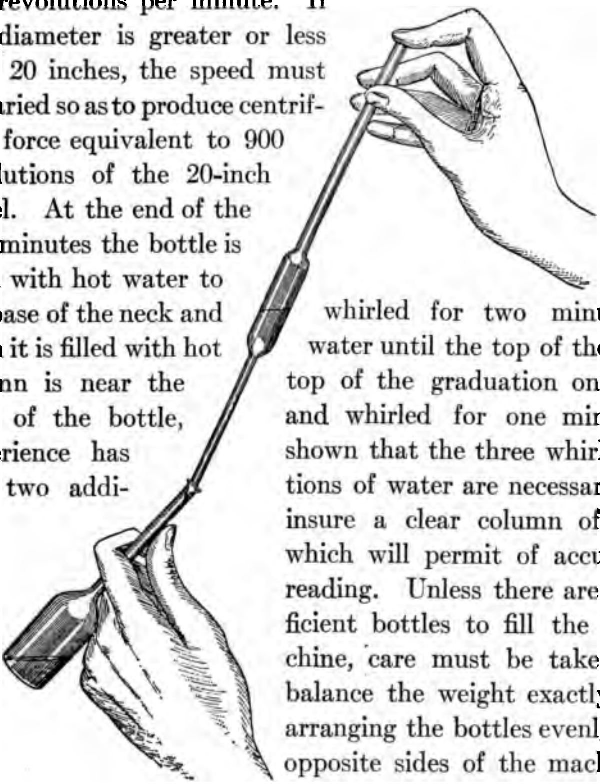

FIG. 166.—Transferring milk from pipette to test bottle.

414. Reading the Fat. The fat must be in a liquid condition in order to make an accurate reading. After the whirling is completed the bottle should be allowed to stand for a few minutes in water at 150° F. The fat should then be read from the lowest to the highest point, the difference between the bottom and top readings being the percentage of fat in the milk. Each of the smaller divisions is two tenths of a per cent, and as the fat can be read between the divisions, it is possible to state the percentage to one tenth of one per cent. Readings may be more accurately made by the use of a small pair of dividers. The dividers should be so set that the points are at the extreme top and bottom of the fat column (Fig. 167); then if the lower point of the divider is placed at the zero mark, the upper part will indicate the percentage of fat. Accurately graduated test bottles and pipettes can usually be easily obtained. In some states, however, all the glassware for the Babcock test is required by law to be verified by the experiment station or other agency designated by the legislature of the state.

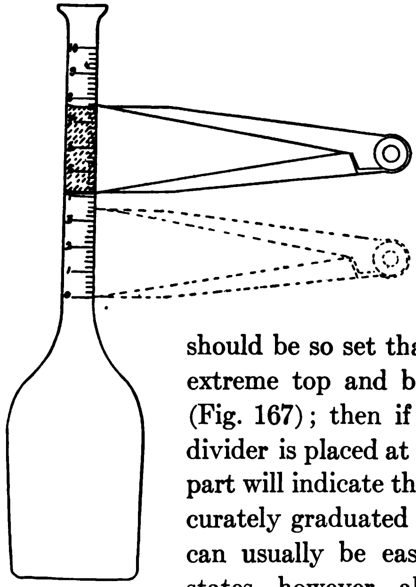
FIG. 167. — Reading the amount of butter fat.

415. Testing Skim Milk and Buttermilk. The skim milk from the separator should frequently be tested to make sure that the separator is thoroughly skimming the milk. Likewise the buttermilk is tested to guard against losses of fat due to incomplete churning. The tests for these products are made in the same way as for whole milk, except that a special double-neck bottle (Fig. 168) is used. The milk and acid are poured into the bottle through the larger side neck. The smaller neck in the center of the bottle is graduated to read 0.05 per cent and may be estimated between the marks to 0.01 per cent. If the separator is working properly, the test will not show more than 0.05 per cent of fat in the skim milk, and good churning should leave not to exceed 0.2 to 0.3 per cent of fat in the buttermilk.

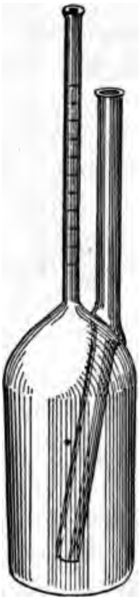

FIG. 168. — Test bottle for skim milk and buttermilk.

416. Testing Cream. It is much more difficult to make an accurate test of butter fat in cream than in milk. Cream varies in fat content from about 12 per cent to over 40 per cent. As the test bottle for milk reads to 10 per cent only, it is necessary to have a special bottle for cream (Fig. 169). The bottle in common use has a much wider neck than the milk bottle and is graduated to 30 per cent. The cream cannot be accurately measured, since the pipette will not deliver 18 grams of cream. The larger the percentage of fat, the lower the weight of cream delivered by the pipette. To make an accurate test it is necessary to weigh the cream instead of measuring it. This may be done on any balance, but when a great many tests are to be made the specially designed balance (Fig. 170) will be found convenient. The bottle is first balanced; then an 18-gram weight is placed on the side opposite to the bottle, and cream

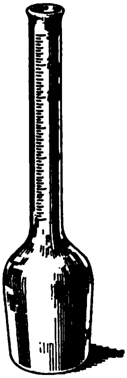

FIG. 169. — Bottle for testing cream.

is slowly run into the bottle until it just balances the weight. The test is then made in the same manner as for milk. If the cream runs over 30 per cent fat, 9 grams are used and nine cubic centimeters of water are added to the cream in the bottle before the acid is introduced. The reading is then doubled to give the correct percentage of fat.

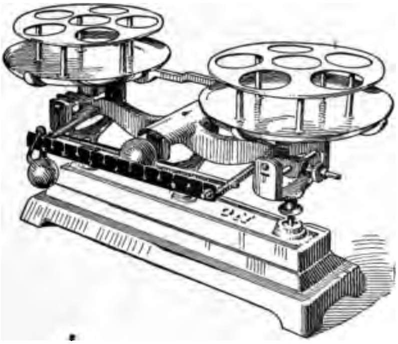

FIG. 170. — Scales used in cream test.

A fairly accurate test of the fat in cheese may be made by weighing out 9 grams of the sample and placing it in the cream bottle with 17.6 cubic centimeters of warm water. The cheese should be allowed to soak for an hour or more, after which the acid should be added, and the test should be conducted in the usual way. To obtain the percentage of fat in the cheese the reading should be doubled.

417. Composite Tests. Where large herds of cows are to be tested, or in the case of creameries with a large number of patrons, the daily testing of the milk is expensive and time consuming. In such cases very good results may be obtained by use of the composite test. A small sample from each milking is placed in a bottle with a preservative such as potassium bichromate or mercuric chloride (corrosive sublimate), and a test of the composite sample is made at the end of a week or ten days. To make such a composite sample represent the true average of the various milkings, the amount added to the bottle each time should be proportional to the weight of the milk at that milking. The entire sample should be well mixed after each addition of milk. The test is made as described for ordinary milk. Various devices have been invented for taking proportional samples of the milk. The desired result may be obtained also by taking 2 cubic centimeters of the milk for the composite sample for each pound in weight of the milk at each milking.

418. Use of the Lactometer. By the use of the lactometer and the Babcock test it is possible to make a fairly accurate estimate of the other solids in the milk in addition to the fat. In the laboratory, the creamery, and the cheese factory it is customary to speak of all the solids of milk exclusive of the fat as *solids not fat* (S. N. F.).

The lactometer is an apparatus for determining the specific gravity of the milk. It consists of a glass spindle with a long slim neck, which will float in milk or other liquid. The lower the specific gravity of the liquid, the deeper the spindle sinks in it, and the heavier the liquid, the higher the spindle floats. All the solids found in milk with the exception of the fat are heavier than water and, therefore, increase the specific gravity of the milk. Since butter fat, on the other hand, is lighter than water (specific gravity about 0.9), the presence of fat counteracts the increase of specific gravity due to the other solids. Milk varies in specific gravity from 1.029 to 1.034. The lactometer most commonly used is the Quevenne lactometer (Fig. 171). This instrument is so graduated that the figures on the neck represent the second and third decimal places of the specific gravity; that is, 32 on the lactometer is equivalent to a specific gravity of 1.032. These figures of the scale on the neck of the instrument are called for convenience *lactometer degrees*. The lactometer reading is made by placing the milk in a tall cylinder and floating the lactometer in it. When the lactometer has come to rest, the reading is taken at the level of the milk's surface. The lactometer gives the correct reading only when the temperature of the milk is 60° F. If it is not at that temperature, 0.1 should be added to the lactometer reading for each degree of temperature above 60°, or the same number should be subtracted for each degree below. In no case should readings be made if the temperature varies more than ten degrees from the normal (60° F.).

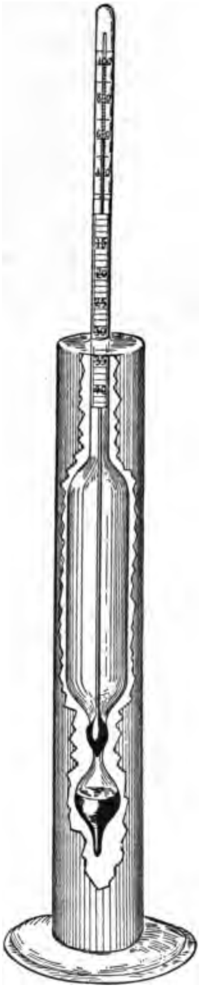

FIG. 171. — Lactometer.

It has been found by experiment that the percentage of solids not fat in a milk is equal to one fourth of the lactometer reading plus two tenths of the percentage of fat. This statement is briefly expressed in what is known as the Babcock formula for solids not fat:

$$\text{S.N.F.} = \frac{L}{4} + .2f.$$

In the formula L stands for the lactometer reading and f for the percentage of fat. For example, if the lactometer reading is 34 and the fat is 3.5 per cent, then

$$\text{S.N.F.} = \tfrac{34}{4} + 0.7 = 9.2.$$

The sample of milk, therefore, contains 9.2 per cent of solids not fat.

The variations in the solids other than fat in milk are not very great even when the different breeds of cows are considered. Adding water to milk decreases the percentage of solids not fat. The above formula, therefore, may be used to detect watered milk. Many states have established by law a standard for fat and total solids in milk. A common legal requirement is a minimum of 3.25 per cent of fat and 9 per cent solids not fat. If the solids not fat in a milk are below 9 per cent, the milk is assumed to be watered. If the fat falls below 3.25 per cent, the milk is declared to be skimmed. By the use of the Babcock test and the lactometer it is possible, therefore, to determine whether the milk has been skimmed or watered, or both skimmed and watered.

419. Acid Test. It is often desirable to know the amount of acid in a sample of milk or cream. This can easily be determined by the use of the Farrington alkali tablets. These tablets consist of a carefully weighed quantity of sodium carbonate mixed with a little phenolphthalein indicator (138). To make the test, dissolve 5 tablets in 97 cubic centimeters of water. Measure 17.6 cc. of the milk into a white cup and add the alkali solution slowly, stirring the milk thoroughly after each addition. The acid in the milk will at first destroy the color of the indicator. The alkali solution is added until a pink color is imparted to the milk which does not immediately disappear upon stirring. Each cubic centimeter of the alkali used is equivalent to 0.01 per cent of lactic acid; for example, if 35 cc. of the alkali solution are used the acidity of the milk is said to be 0.35 per cent.

FIG. 172. — **Graduated cylinder and white cup used for a test of milk or cream.**

EXERCISES

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CHAPTER XLIV

LEAVENING AGENTS

420. Wheat Bread. Bread of some kind has been used by mankind from the earliest times. During the earlier periods it consisted chiefly of powdered meal and water, baked in the sun or on hot stones. This kind of bread has the same characteristics as modern hard-tack or hoe-cake, so far as digestibility is concerned. It has great density, is difficult to masticate, and presents but little surface to the action of the digestive juices. Such bread is said to be *unleavened.*

Very early in the history of the human race leavened bread seems to have been used. Experience must have taught the semicivilized tribes that a light and porous loaf was more palatable and more digestible than a dense one, for nearly all peoples now have some method of "raising" bread and making it light.

421. To Make a Light Loaf. The incorporation into the dough of bubbles of some kind of gas is necessary to make a light loaf. The possibility of a light, porous loaf of wheat bread is due to the gluten of the wheat (319). This substance, when moistened, forms a sticky elastic mass, which incloses the gas bubbles and prevents their escape. When the bread is heated in baking, the gluten loses its elasticity, and becomes stiffened, and retains its light porous condition. Three general methods are used to introduce the gas, namely, mechanical, biological, and chemical methods.

The best known example of the mechanical methods is the "beaten biscuit" of the South. In this case air is introduced

into the dough by long-continued beating, which makes it somewhat porous. When the dough is heated in the oven, the air expands and increases the porosity of the product.

The so-called aërated bread is made by mixing the dough with carbonated water such as is used in soda-water fountains. The mixer is so constructed that the carbon dioxide is kept under high pressure during the mixing. When the mixing is completed, the pressure is released and the expansion of the gas makes the dough light, as does the further expansion of the gas when the dough is placed in the oven.

The use of eggs for making light cakes, such as sponge cake and angel cake, is another example of the mechanical method. The air which is beaten into the egg albumin supplies the bubbles of gas needed to make the cake light.

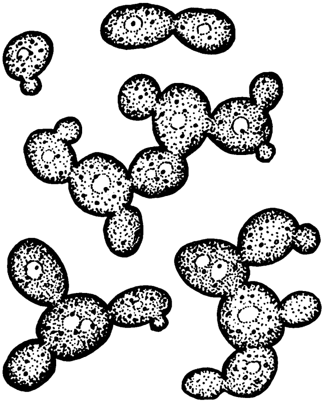

FIG. 173. — Yeast plants.

422. Use of Yeast in Making Bread Light. Yeast is the substance most commonly used to make bread light. Condensed yeast consists of a mass of microscopic plants (Fig. 173) which grow or multiply rapidly in the moist dough or *sponge*, especially if some sugar has been added. During its growth the yeast plant secretes enzymes which bring about the fermentation of the sugars with the formation of alcohol and carbon dioxide (362):

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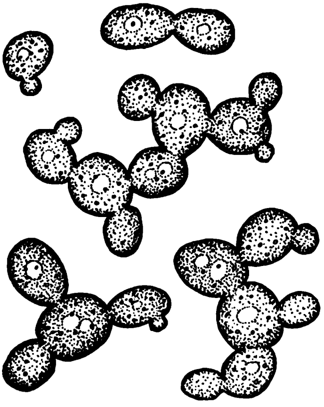

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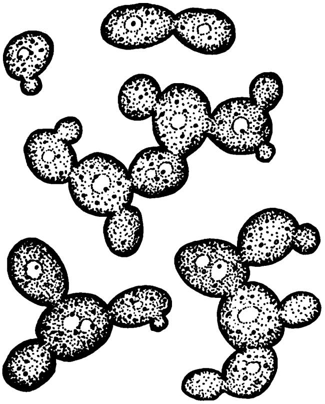

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into the dough by long-continued beating, which makes it somewhat porous. When the dough is heated in the oven, the air expands and increases the porosity of the product.

The so-called aërated bread is made by mixing the dough with carbonated water such as is used in soda-water fountains. The mixer is so constructed that the carbon dioxide is kept under high pressure during the mixing. When the mixing is completed, the pressure is released and the expansion of the gas makes the dough light, as does the further expansion of the gas when the dough is placed in the oven.

The use of eggs for making light cakes, such as sponge cake and angel cake, is another example of the mechanical method. The air which is beaten into the egg albumin supplies the bubbles of gas needed to make the cake light.

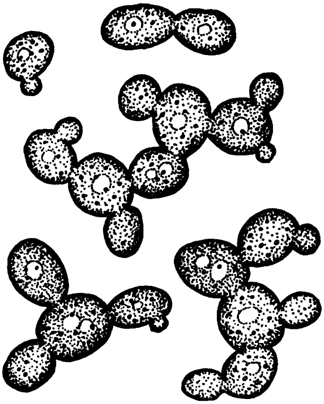

FIG. 173. — Yeast plants.

422. Use of Yeast in Making Bread Light. Yeast is the substance most commonly used to make bread light. Condensed yeast consists of a mass of microscopic plants (Fig. 173) which grow or multiply rapidly in the moist dough or *sponge*, especially if some sugar has been added. During its growth the yeast plant secretes enzymes which bring about the fermentation of the sugars with the formation of alcohol and carbon dioxide (362):

$$\underset{\text{glucose}}{C_6H_{12}O_6} \rightarrow 2\ \underset{\text{alcohol}}{C_2H_5OH} + 2\ \underset{\text{carbon dioxide}}{CO_2}.$$

When yeast is added to the dough, and the mixture is put in a warm place, the yeast grows and carbon dioxide is produced. The gas permeates the dough, but cannot escape because of the tenacious character of the mass. As a result the dough rises, that is, it swells up because of the bubbles of gas that have been formed within it.

After the dough has risen, it is kneaded for the purpose of breaking up the large bubbles and distributing the gas evenly throughout the mixture. When the dough is shaped into loaves, the fermentation continues because the yeast is still growing, and the further evolution of carbon dioxide causes the dough to rise again.

FIG. 174. — **Cross sections of light loaves of bread. The variation in size is due to differences in the quality of the gluten in the two flours.**

When the loaves are placed in the oven, the first action of the heat expands the bubbles of carbon dioxide and the loaf is made more porous and light. The yeast is killed, fermentation stops, and the alcohol and part of the water are vapor-

ized and driven off. At the same time the starch grains are ruptured by the heat, and the gluten is stiffened and loses its elasticity. The heat of the oven (350° to 400° F.) converts part of the starch on the outer edge of the loaf into dextrin (312), which is one reason why the crust of the bread is sweeter than the crumb. The interior of the loaf never gets much above 212° F., because the water in the interior of the loaf is being continually changed into steam.

423. Salt-rising Bread. This is another example of a bread which is made light by a biological process. To make salt-rising bread, corn meal and a little baking soda are mixed with warm milk. After some hours this batter begins to ferment and carbon dioxide is produced. Dough is made as for ordinary bread and this mixture is used in the place of yeast. The process of handling is from this point the same as that described for yeast bread. For a long time the cause of the fermentation in salt-rising bread was not understood. It has been discovered recently, however, that it is due to certain gas-forming bacteria that are found associated with the corn meal.

FIG. 175. — Making bread in the Orient. Kneading dough with leaven.

The ancient method of preserving a little of the dough from each batch of bread to act as the leaven for the next is still followed in some countries. This small bit of old dough carries a sufficient number of yeast plants to start the fermentation when it is mixed with a new lot of dough, and thus "a little leaven leaveneth the whole lump."

424. Ammonium Bicarbonate. The use of ammonium bicarbonate is the simplest of the chemical methods of raising bread. The ammonium salt is dissolved in the water used in making the dough. When the mass is heated, the bicarbonate decomposes into gaseous products:

$$NH_4HCO_3 \rightarrow NH_3 + H_2O + CO_2.$$

The evolution of ammonia and carbon dioxide and their expansion when heated make the bread porous and light. In skillful hands this method is very successful and has the advantage of leaving no solid residue. The ammonia is driven off by the heat of baking.

425. Sour Milk and Soda. These leavening agents have long been used. The lactic acid in the sour milk reacts with sodium bicarbonate, also called baking soda, or saleratus, and liberates carbon dioxide:

$$\underset{\text{lactic acid}}{H{\cdot}C_3H_5O_3} + NaHCO_3 \rightarrow \underset{\text{sodium lactate}}{Na{\cdot}C_3H_5O_3} + H_2O + CO_2.$$

When properly followed, this method of leavening is very successful, but it has the disadvantage that it is sometimes difficult to adjust the proportions of the sour milk and soda. If too much sour milk is used, the product has a sour taste. If too much soda is used, the excess of bicarbonate is changed by the heat of the oven into the normal carbonate, which gives a soapy taste to the bread:

$$2\ NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O.$$

Occasionally baking soda and hydrochloric acid are used as leavening agents. In this case the soda is mixed with the flour and the hydrochloric acid is added to the water which is used in mixing the dough:

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2.$$

426. Baking Powders. The difficulty of measuring the exact amounts of acid and bicarbonate when the above methods are used has led to the manufacture of commercial baking powders. In baking powders an acid salt is used instead of the free acid.

Tartrate baking powder is composed of baking soda and cream of tartar. These substances do not react with each other while dry, but when moistened the following reaction takes place:

$$NaHCO_3 + KH \cdot C_4H_4O_6 \rightarrow KNa \cdot C_4H_4O_6 + H_2O + CO_2.$$

Phosphate baking powders consist of baking soda and calcium mono-phosphate, the same compound that occurs in the fertilizers known as acid phosphate or superphosphates. The reaction in this case is as follows:

$$2\ NaHCO_3 + CaH_4(PO_4)_2 \rightarrow CaHPO_4 + Na_2HPO_4 + 2\ H_2O + 2\ CO_2.$$

Alum baking powders are made from baking soda and ammonium alum. The reaction here is a deep-seated one:

$$2\ NH_4Al(SO_4)_2 + 6\ NaHCO_3 \rightarrow 2\ Al(OH)_3 + 3\ Na_2SO_4 + (NH_4)_2SO_4 + 6\ CO_2.$$

Starch is added in the manufacture of all baking powders to keep the materials dry, and to coat each particle of the acid salt and the carbonate in order to prevent their acting upon each other while stored. From 20 to 25 per cent of starch is sufficient for this purpose.

427. Healthfulness of Baking Powders. It will be seen from the above equations that one or more products of chemical reaction remain in the food after baking. The possible harmful action of these substances has been the subject of much discussion, but there are few reliable experimental data on the subject. There is no reason to believe that the sodium lactate produced when soda and sour milk are used has any harmful effects. Cream of tartar and soda produce Rochelle salt, which is a laxative, but which probably has no effect in the small quantities consumed with the food. The same statement may be made regarding the sodium phosphate produced by the phosphate baking powders. More has been said against the alum baking powders than any of the others, and many physicians believe that the residue of aluminum hydroxide left by them is injurious to health. Since nothing is to be gained by their use except a very small saving in expense, it would be well to avoid these powders, so long as there is a question about it.

428. Homemade Baking Powder. A baking powder can be prepared at home from the following ingredients:

Cream of tartar, dried	1 pound
Baking soda	½ pound
Starch	½ pound

Thoroughly dry the starch and cream of tartar in a warm (not hot) oven. Divide the starch into two parts, and mix the soda with one part, and the cream of tartar with the other. Then mix the whole thoroughly and keep the mixture in cans or bottles in a dry place.

429. Shortening. Pie crust and similar forms of pastry are not leavened. Instead the flour is so treated that when heated it crumbles readily into thin flakes. This is accomplished by the method known by the housewife as shortening, and consists in mixing a fat, such as butter, lard, or oil, with the flour in making the dough. The fat destroys the elasticity of the gluten, making it break off short when worked instead of allowing it to remain tenacious, as it is in ordinary dough. The result of this treatment is a flakiness that has the effect of exposing a large surface to the digestive fluids; but this good effect is counteracted by the presence of the large amount of fat, which produces a greasy surface that interferes with the action of the digestive juices upon the proteins and carbohydrates. Because of its expansion under the heat of the oven, the air folded into the dough is probably another factor in making the pastry light and flaky.

EXERCISES

Ex. **261.** Is the use of wheat bread of modern origin? How was the first bread probably made? Why is a light loaf desirable? What is necessary to make a light loaf? What is the function of the gluten in bread making? What change takes place in the gluten when the bread is baked?

Ex. **262.** What is the principle underlying the making of beaten biscuits? How is aërated bread made? What makes it light? Explain how eggs make possible the production of a light cake such as sponge cake.

Ex. **263.** Explain how the use of yeast makes a light loaf possible. Devise an experiment to show that carbon dioxide is given off by the sponge in bread making. Explain the formation of the numerous holes in a loaf of bread. Why does the baked loaf retain its shape when the dough will not?

Ex. **264.** What makes the bread light in the case of salt-rising bread? Why is the corn meal added? Explain the ancient method of saving the leaven from one baking to another.

Ex. **265.** When ammonium carbonate is used to raise bread, what chemical change takes place? What advantage has ammonium carbonate over other baking powders? Explain the use of baking soda with sour milk. Add some baking soda to sour milk at home and prove that carbon dioxide is given off. What is the chief difficulty in using these materials? What happens if too much soda is used?

Ex. **266.** Moisten samples of baking powder with water and determine whether carbon dioxide is evolved. What are the three general types of baking powders on the market? Obtain several samples of baking powder and test to see if they are tartrate, phosphate, or alum powders, as follows:

(*a*) Test for phosphate by burning the powder to destroy the starch, and then heat with a little nitric acid and add ammonium molybdate reagent.

(*b*) Test for alum. Dissolve in water and filter. To the filtrate add a little hydrochloric acid and some barium chloride solution. A white precipitate indicates alum. (*Note.* This is really a test for sulphates, but as alum is the only sulphate used in baking powders it serves the purpose.)

(*c*) There is no easy test for tartrate, but if the powder is not a phosphate nor an alum powder, it is undoubtedly made with tartrate.

Ex. **267.** What is the residue left in the food when sour milk and soda are used? When cream of tartar and soda are used? When a phosphate powder is used? In case of an alum powder? Is any objection raised to any of these powders? What type of baking powder do you use at home?

Ex. **268.** Have the class make a baking powder according to the directions in paragraph 428. Divide among the members of the class and have it tested in their homes. Call for reports.

Ex. **269.** (Teacher) From a cupful of flour prepare gluten as in exercise 197. Note the elasticity of the gluten. Now work thoroughly into the gluten a good-sized piece of lard. What effect does the fat have on the elasticity of the gluten? What is meant by shortening? How does it make pastry light and flaky? Why is pastry hard to digest?

CHAPTER XLV

FOOD PRESERVATION, ANTISEPTICS, AND DISINFECTANTS

430. Why Foods Spoil. The fermentation and decay of foods, which render them unfit for consumption, are due to the growth of molds and microörganisms such as yeast and bacteria (Fig. 176). These organisms act upon the carbohydrates and produce alcohol, carbon dioxide, and various organic acids. Some of the bacteria cause the decay or putrefaction of the proteins, with the production of the foul odors so noticeable in decaying meat and other high protein foods. Some of these organisms produce from the proteins poisonous substances which are known as ptomaines (326). Without some method of preserving food the human diet would at least lack the variety that is now possible. Molds, yeast, and bacteria are present everywhere in the air, especially when the air is dust laden. While foods should be protected from dust and dirt, which always contain bacteria, it is practically impossible to prevent entirely their entrance into the food ma-

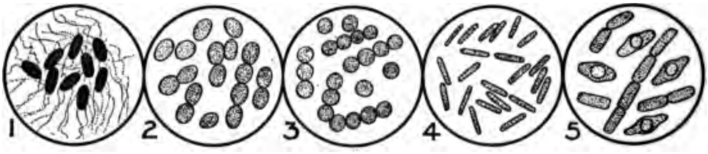

FIG. 176. — **Bacteria.** 1. Typhoid fever. 2–5. Forms of bacteria found in milk.

terials. Hence if food is to be preserved, some method must be used to destroy these organisms or to prevent their growth.

431. Preservation by Drying. The conditions necessary for the growth of the organisms that cause foods to spoil are the presence of moisture and warmth. Materials that are very dry, such as flour and crackers, will keep indefinitely if not allowed to become damp. The drying of meats, fruits, and vegetables has been practiced from early times. Formerly these food products were all dried in the sun; but now many devices are in use for the artificial drying of these materials.

FIG. 177. — The preservation of foods by drying.

432. Refrigeration. A low temperature retards the growth of the organisms that produce fermentation and decay, and consequently the lower the temperature at which foods are held the longer they will keep. The cold storage business, which has grown to enormous proportions, is based upon this principle. Food substances that can be frozen without injury can be kept almost indefinitely in the frozen state. Since cold and freezing do not kill the bacteria outright but merely prevent their growth, perishable substances when removed from cold storage will, of course, spoil the same as fresh materials. The temperature produced in the home ice box or refrigerator is not low enough to retard the decomposition of food for very long periods.

433. Preserving in Strong Solutions. Yeasts and bacteria cannot grow in concentrated solutions. When the solution contains more than 20 per cent of solids, the osmotic action of the solution is so great that it extracts water from the living cell and causes the protoplasm to collapse, resulting in the death of the cell. The method of preserving meats and vegetables by placing them in strong brine depends upon this fact. If the brine is too weak, the organism can grow and the food spoils. Dry salting of meats is another instance of the same method of preservation. The salt rubbed into the meat dissolves in the moisture present and forms a brine which is strong enough to prevent the growth of bacteria.

Approximately fifty per cent of jelly and jam consists of the sugar added in their manufacture. Sugar is a food for bacteria and yeast when in dilute solution, but in the concentration used in jellies and jams it destroys these organisms by its great osmotic action. The heating incident to preparing these foods also assists in destroying these organisms.

434. Preservation by Heat. The organisms that cause fermentation and decay in foods are killed at the temperature of boiling water if the heating is continued for some time. Foods that are boiled or otherwise heated to this temperature will, therefore, keep indefinitely if properly protected from further infection. The canned vegetables which occupy such a large place in the modern dietary are preserved by this method. The vegetables are placed in cans and heated and the cans are sealed while very hot. The heating is commonly done under pressure to make the temperature still higher and the destruction of the organisms more certain. If the ends of the can bulge outward or if gas escapes under pressure when the can is punctured it is a sign that the food in the can has fermented. Such goods should be rejected.

Bottled fruits and vegetables prepared at home depend upon the same principle for their preservation. If the bottles and the tops and the rubbers are sterilized and the foods are completely sterilized by cooking, they will keep indefinitely. If any of the germs are undestroyed, fermentation will subsequently take place. Some bacteria form spores with very thick cell walls which are resistant to heat. Unless all these are killed they may later germinate and cause the food to ferment. They may be killed by long-continued heating, or by the very high temperature produced under pressure. As too long a period of boiling injures the quality of some of the food products, the method of intermittent heating is sometimes used. In this method the food is heated for short periods at three different times, a day or two apart. The spores germinate in the intervals between the heatings and thus may easily be killed. Small pressure heaters suitable for use in the home are now on the market, and their use makes the canning process more reliable.

A modification of the canning procedure known as the *cold pack method* (Fig. 178) is growing in favor and is largely used by the members of the *girls' canning clubs.* In this method the material is first *blanched,* that is, cooked for a short time in boiling water or steam. It is quickly dipped into cold water, and afterward packed into hot jars which are then filled with boiling water, immersed in hot water, and boiled for periods of from ten minutes to three hours, depending upon the material that is being canned.

It has been stated that temperatures of 140° to 150° F. will kill most of the germs and delay the spoiling of foods. About the only practical application is the pasteurizing of milk.

435. Chemical Preservatives. Sometimes chemicals are added to foods to prevent their fermentation. The chemicals most commonly used are borax, boric acid, benzoic acid, sodium benzoate, salicylic acid, sodium salicylate, formaldehyde, sulphur dioxide, and sodium sulphite. The use of some of these substances is illegal, while others may be legally used if their presence is indicated on the package. The harmfulness of some of these preservatives is a matter of dispute; but as their use is unnecessary, it is best to avoid all canned or bottled products containing any preservative. The use of these chemicals makes it possible to can poor products under careless and unsanitary conditions. With good vegetables or fruits and clean factory conditions the use of chemical preservatives is absolutely unnecessary, and it seems reasonable to believe that any chemical that will prevent the growth of bacteria will also interfere with digestion. The

FIG. 178. — **Preserving by the cold pack canning method.**

use of formaldehyde in milk is especially pernicious, as milk is so largely used in feeding babies and young children. A sample of milk or moist food product which keeps a long time when exposed to the air probably contains a chemical preservative. Such products should not be used as foods.

436. Antiseptics are substances that check the growth of bacteria. They may not kill all the bacteria present, however. Some of the same substances that are used as food preservatives are also employed as antiseptics. Most of the antiseptic mouth washes, such as listerine, have borax and boric acid as their basis, associated usually with thymol or some other aromatic antiseptic. A solution of boric acid is often used also as an antiseptic eye wash. Various antiseptics are used also as a wash or dressing for wounds in order to destroy any bacteria which may have found their way into the wounds. Such bacteria as the tetanus germ that produces lockjaw are often present in dirt and may be introduced into a wound. Hydrogen peroxide, a weak solution of carbolic acid, listerine, or even alcohol may be employed as antiseptic applications in such cases. Every cut or other break in the skin should be immediately washed with one of the above solutions, especially if any dirt is known to have entered the wound. Physicians often dress wounds with iodoform, which is another antiseptic.

437. Disinfectants, or germicides, are substances that will kill all bacteria. Some of these materials are antiseptics in dilute solutions and disinfectants in greater concentration. Disinfectants are used to destroy bacteria and thereby remove the danger of infection. A solution of mercuric chloride (corrosive sublimate) is quite commonly used as a disinfectant, as is also strong carbolic acid. Bleaching powder, or chloride of lime (124), has strong germicidal properties because of the chlorine liberated from it. All articles used in cases of contagion should be immersed in a disinfectant solution before being removed from the sick room. Heat is one of the best disinfectants, and articles that will stand boiling in water can best be sterilized in that way.

Sulphur dioxide has long been used for disinfecting bedrooms after cases of contagious diseases. It is often formed by burning sulphur in the room; or the liquid sulphur dioxide, which may be purchased in small metal cylinders, is used. Both chlorine and sulphur dioxide are bleaching agents, and their use as household disinfectants is largely superseded by the use of formaldehyde, which is a powerful germicide and has no bleaching action (287). Specially devised machines for manufacturing formaldehyde on the ground are used when large areas are to be fumigated.

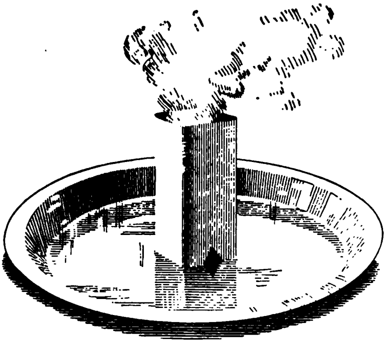

FIG. 179. — A formacone candle.

For use in disinfecting on a smaller scale, the solid substance sold under such names as *formacone* or *paraform* is convenient. This is a condensation product of formaldehyde. Under certain conditions several molecules of formaldehyde can be made to combine to form the white solid substance mentioned above, which is known to the chemist as paraformaldehyde. This substance is made into tablets or candles and when heated or boiled with water it is changed back into gaseous formaldehyde, and thus provides an easy and efficacious method of disinfecting a room after sickness.

EXERCISES

Ex. **270.** What causes foods to spoil? How generally are molds and bacteria distributed? How does a knowledge of food preservation affect our dietary? What conditions are necessary for the growth of molds and bacteria? Why do not perfectly dry foods spoil? What can you say of drying as a method of food preservation? What effect has cold on the keeping of food? Does freezing kill the bacteria outright?

Ex. **271.** How does a strong brine prevent the fermentation of foods? If the brine is too weak what happens? Explain the preservation of meat by dry salting, Why are jellies and jams so easily kept? Would fruits keep if only a little sugar was added to them?

Ex. **272.** Explain the theory of keeping fruits and vegetables by canning. Why do factories heat them under pressure? Why should canned goods be discarded if the can bulges outward? Why are vegetables sometimes heated on three different days? Does anyone in your neighborhood use a pressure heater in canning fruits and vegetables? What is meant by the *cold pack method?* Write to your Agricultural College for bulletins describing the cold pack method and try canning some kind of vegetable by this method.

Ex. **273.** What chemicals are sometimes used as food preservatives? What do you think of the policy of using these chemicals? Are they necessary with good materials? Examine the cans and bottles that come into your home and see if any of them state that a preservative is used.

Ex. **274.** To a pint of milk add two or three drops of formaldehyde. Set the bottle aside and see how long it takes the milk to sour. Why should milk be regarded with suspicion if it keeps too long? Test suspected milk for formaldehyde as follows: Fill a test tube one third full of the milk. Hold the tube in an inclined position and carefully pour down the side of the tube a little sulphuric acid to which has been added a drop of ferric chloride solution. If formaldehyde is present a violet color forms where the acid and milk came into contact.

Ex. **275.** What is meant by antiseptics? Of what are most of the antiseptic washes made? Why is it dangerous to get dirt into a wound? How should cuts be treated? Name some of the more common antiseptics. What is a disinfectant? In what way does it differ from an antiseptic? Name some disinfectants. Why is formaldehyde now used in place of sulphur dioxide and chlorine?

CHAPTER XLVI

TEXTILES, DYEING, AND BLEACHING

438. All the fabrics ordinarily employed in making clothing are composed of one or more of the four commonly used fibers: cotton, linen, wool, or silk. Of these fibers, cotton and linen are of vegetable origin, while wool and silk are derived from animal sources.

439. Cotton. Cotton fibers are the seed hairs of the cotton plant. Each seed hair is a long, tubular, single plant cell, which during growth is full of protoplasm. As the seed ripens, the protoplasm disappears, the tube collapses, and the hair becomes twisted into a spiral as shown in Figure 180.

Cotton Wool Silk

Fig. 180. — **Textile fibers.**

Cotton when pure consists almost wholly of cellulose. It is not readily attacked by alkalies, but is easily destroyed by acids (313). Cotton fabrics are used in larger quantities than those made from any other fiber.

440. Mercerized Cotton. When cotton is treated with a concentrated solution of sodium hydroxide, it contracts to about three fourths of its original length, and is converted into a new substance called alkali cellulose. If this material is now stretched to the original length of the cotton and then thoroughly washed and dried, the fiber takes on a silky sheen. *Mercerized cotton* is stronger than ordinary cotton and has a greater affinity for dyes.

Cotton is treated also in a number of ways to give the fiber the appearance of silk. Most of the artificial silks, such as near silk, chardonnet silk, and viscose silk, are manufactured from cotton fiber by special treatments.

441. Linen also is almost pure cellulose. It is made from the fibers of the straw of the flax plant. The straw is placed in stagnant water where it partially decays, this decay making it possible to separate the fibers from the other parts of the plant by mechanical means. This process is known as retting. Natural retting is sometimes replaced by so-called chemical retting, a process in which the fibers are separated by means of dilute acids instead of by the slower process of decay. The fiber of linen is longer and stronger than that of cotton; it has more luster, and is a better conductor of heat. Linen is also more readily destroyed by strong alkalies and by chlorine and other oxidizing agents.

442. Wool is the hair of the sheep or goat. There are several varieties of wool, the kind depending upon the animal from which it is obtained. Cashmere comes from the Tibet goat, mohair from the Angora goat, and alpaca from the llama. Not all dress goods bearing these names come from the proper sources, however, since the peculiar characteristic of each can be imitated in common wool.

Wool is composed of nitrogenous substances containing sulphur and when burned gives off the peculiar odor of burned hair. All wool fibers have an outer layer of flat cells the edges of which project outward, making a saw-tooth appearance under the microscope (Fig. 180). When the wool fibers are beaten together, this peculiar structure causes them to interlock, and thus makes possible the manufacture of such materials as felt. Wool is readily attacked by alkalies; even dilute solutions of sodium hydroxide cause it to dissolve. On the other hand, acids affect cotton much more readily than they do wool. Much of the cloth that is sold as woolen goods is a mixture of wool and cotton fibers. It is a simple matter to detect such a mixture; for if the cloth is placed in a dilute solution of sodium hydroxide, the wool will dissolve, while the cotton will remain. An all wool cloth will dissolve completely (Fig. 181).

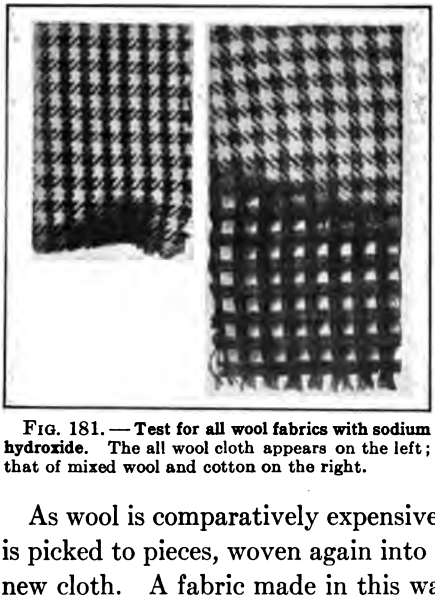

FIG. 181. — **Test for all wool fabrics with sodium hydroxide.** The all wool cloth appears on the left; that of mixed wool and cotton on the right.

As wool is comparatively expensive, old worn woolen cloth is picked to pieces, woven again into yarn, and used to make new cloth. A fabric made in this way is called shoddy. It has a short fiber and, consequently, is weak and will not wear as well as new wool.

443. Silk is obtained from the cocoon of the silkworm. The cocoons are heated to 70° C. to kill the worms and then are washed in warm water to soften the silk glue so that the fibers may be reeled. The fibers from several cocoons are twisted together in the reeling process and thus a silk thread is obtained. Silk thread has a beautiful luster and a high tensile strength. It is a nitrogenous substance containing no sulphur, differing in this respect from wool. Silk is somewhat more resistant than wool to the action of alkalies, but is more readily attacked by acid and is very sensitive to the action of chlorine and other oxidizing agents.

444. Dyeing. The animal and plant fibers differ greatly in their behavior toward dyestuffs. Many of the substances used to dye cloth are either acid or basic in character. The animal fibers resemble the proteins in chemical behavior and like the proteins have the power of uniting chemically with the acid or basic dye and forming a colored compound that is insoluble in water. It is, therefore, a comparatively easy matter to dye silk or woolen goods.

Cotton and linen, on the other hand, are largely cellulose, which has no chemical affinity for the dyes. Cellulose is stained by the dyes, but the color is not fast. Cotton and linen are dyed by first introducing into the fiber an insoluble substance that will unite with the dye and hold the color fast. A common method of procedure is first to immerse the cloth in a solution of aluminum chloride or sulphate, or a bath of ordinary alum; then when the cloth is placed in ammonia water, the following reaction takes place:

$$AlCl_3 + 3\ NH_4OH \rightarrow Al(OH)_3 + 3\ NH_4Cl.$$

The aluminum hydroxide ($Al(OH)_3$) is insoluble and adheres to the fibers of the cloth. If the cloth is now immersed in the dye solution, the aluminum hydroxide unites with the dye and holds the color fast. Materials used in this way to fasten the color in fabrics are called mordants. Salts of aluminum are the more common mordants, although salts of tin, iron, and chromium are largely used, as well as tannic acid, the latter especially for the basic dyes.

445. Direct Dyes. A few years ago it was thought to be impossible to dye cotton and linen without the use of a mordant to hold the color to the fiber. Recently, however, a number of dyes have been discovered that adhere to the cotton and linen, and some of them possess a satisfactory permanence. These direct dyes are used extensively in dyeing mixed goods consisting of cotton and wool or cotton and silk. The colors are not so fast as those used with a mordant, and they are much more readily affected by strong soaps and alkalies.

446. Bleaching. The differences in the chemical composition of the various textile fibers and of the coloring materials to be destroyed make it impossible to use a single method for bleaching cotton, linen, wool, and silk.

Cellulose is capable of withstanding the action of chlorine, as well as that of weak acids and alkalies. Cotton, therefore, is almost universally bleached by means of chlorine derived from bleaching powder (124). The chlorine is liberated from the bleaching powder by a weak solution of sulphuric acid, and the chlorine reacts with water to liberate oxygen as follows:

$$H_2O + 2\,Cl \rightarrow 2\,HCl + O.$$

The bleaching is really due to the destruction of the coloring matter by the nascent oxygen (81) and is, therefore, an oxidation process.

Linen fibers are more sensitive to the action of chlorine than are those derived from cotton. Great care, therefore, must be exercised not to weaken unduly the fiber when bleaching with chlorine. The ancient method of bleaching linen, which is still in use in many parts of the Old World, consists in steeping the cloth in a weak alkaline solution and then spreading it on the grass to bleach in the sunlight. The cloth is sprinkled from time to time with water. It is finally dipped into buttermilk and then washed with soap and water.

Wool and silk should never be bleached with chlorine. The bleaching is usually accomplished by means of sulphurous acid or sodium peroxide. When sulphurous acid is used, the goods are moistened and then subjected to the action of sulphur dioxide produced by burning sulphur (63).

When sodium peroxide is used, the goods are first immersed in a dilute solution of sulphuric acid, and then the sodium peroxide is added. This substance reacts with the acid to produce hydrogen peroxide (45) thus:

$$Na_2O_2 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O_2.$$

The hydrogen peroxide readily decomposes, liberating oxygen (81), which is the bleaching agent.

EXERCISES

Ex. 276. Examine cotton fibers under the microscope and make a drawing of them. What is the chemical composition of pure cotton? Pour 20 cc. of sulphuric acid into 10 cc. of water. When the mixture is cool, place a piece of cotton cloth in it for a few minutes. Wash the cloth in water and note what effect the acid has on the strength of the fiber.

Ex. 277. Cover a piece of cotton cloth with a 30 per cent solution of sodium hydroxide for 10 or 15 minutes. Wash the cotton in water, dip it in vinegar, and wash it again. Compare with the untreated cotton. How did the sodium hydroxide affect its strength? Compare this with the effect of the acid in Ex. 276.

Ex. 278. Stretch a small piece of white cotton cloth on a frame and dip it in a 30 per cent solution of sodium hydroxide for 15 minutes. Wash it thoroughly and note the effect of the treatment on the strength and luster of the cotton. Why was the cloth stretched on a frame? What is cotton treated in this way called?

Ex. 279. Examine fibers of linen under the microscope. Note the absence of twist and the presence of "knots" at the junction of the cells. What is the source of linen? What is meant by retting the flax? How does the strength of linen compare with cotton? Which is more resistant to alkalies and chlorine?

Ex. **280**. Burn some woolen cloth and note the odor. Place bits of woolen cloth in a test tube and heat. Hold a piece of moist red litmus paper in the escaping vapor. Result? Hold a piece of lead acetate paper in the vapor. Result? Does wool contain nitrogen and sulphur? What are the sources of wool for cloth making?

Ex. **281**. Examine fiber of wool under the microscope. Note the saw-tooth appearance. Of what advantage is this structure? What is meant by shoddy? Is shoddy as strong as the original woolen cloth?

Ex. **282**. Boil a piece of pure woolen cloth and a piece of mixed wool and cotton in a 10 per cent solution of sodium hydroxide. What happens? How is it possible to distinguish between pure woolen goods and mixed goods? Try the experiment on goods from home.

Ex. **283**. Examine fibers of silk under the microscope and compare with wool. Heat some bits of silk in a test tube and test for ammonia and sulphur. Results? Compare with wool.

Ex. **284**. Dip a piece of cotton cloth in a solution of alum. After a few minutes take out the cotton, wring it out and dip it in a dilute solution of ammonia water. Now dip this piece of cloth and an untreated piece into a solution of logwood and boil it. After a few minutes withdraw the two pieces of cloth and wash them thoroughly in clear water. Is there any difference in the two pieces? Which holds the color better? How do you explain this? What is meant by a mordant? (*Note.* If logwood is not available a strong solution of litmus may be substituted with fairly good results.)

Ex. **285**. Dip pieces of colored cotton cloth in a dilute solution of bleaching powder. Transfer from this solution to a dilute solution of hydrochloric acid. Note the effect on the color. Should bleaching powder be used for wools or silks? How are wools and silks usually bleached?

CHAPTER XLVII

PAINTS AND VARNISHES

447. ALL paints consist of two parts: (1) an opaque solid, and (2) a liquid that holds the solid in suspension while the paint is being spread on a surface, and which causes it to adhere firmly to the substance that it covers. The solid is called the pigment, and the liquid the vehicle.

448. Linseed Oil the Best Vehicle. The most commonly used and the best vehicle for paint making is linseed oil, which is made from flaxseed. The oil is extracted by pressing the ground seed. If the oil is extracted in the cold, it is very light colored, but a more abundant yield of a darker colored oil is obtained if the seeds are subjected to heat while being pressed. Linseed oil, as has been shown (302), absorbs oxygen from the air and is changed into a tough, solid substance. This change takes place readily when the oil is spread in a very thin layer. This property of linseed oil causes it and the pigment to adhere closely to the surface to which the paint is applied. The pigment fills the pores that form in the film of oil and helps to cover the surface more completely than would be done by the oil alone.

449. Boiled Oil and Driers. Linseed oil that has been heated with lead oxide or the oxides of manganese will dry more quickly than the raw oil. Such oil is called *boiled oil.* The more rapid drying is probably due to the catalytic action of the oxides absorbed by the oil in the process of boiling. The paint can also be made to dry more quickly by adding small quantities of so-called driers. These are substances that absorb oxygen from the air and then give it off to the oil, thus hastening the oxidation or drying of the oil itself.

Turpentine is a substance of this kind, and is commonly added to paint when quick drying is desired.

Japan driers are made by boiling a little oil with a large amount of lead or manganese oxides, and dissolving the mixture in turpentine. By adding a small quantity of this drier to the paint it is made to dry more rapidly. Too much drier is objectionable as any such substance really tends to make the paint less durable. Much of what is sold as boiled oil is merely raw linseed oil to which drier has been added.

450. Adulterated Linseed Oil. The comparatively high price of linseed oil has led to its frequent adulteration. Such materials as cottonseed oil or corn oil are substituted in part or in whole for the linseed oil. These substitutes are not drying oils, and paints made with them will not dry. As a consequence such paints are not durable, especially when exposed to the weather. Paints that remain sticky long after application have been mixed with other than linseed oil. The best test is to brush a thin layer on a pane of glass and place it outdoors for forty-eight hours. If the oil has not dried to a hard film with no stickiness, it is not true linseed oil, and is not suitable for use in mixing paints.

451. Common Pigments. Those used to form the body of paints are white lead (250) and zinc oxide (219). White lead is the oldest of the white pigments and is still probably used more than any other. It has good covering quality but has the fault that it rubs off or chalks after a time. It also darkens with age because of the formation of lead sulphide.

Zinc oxide does not cover so well as white lead but is more durable and does not chalk or darken with age. Probably the best white pigment is a mixture of white lead and zinc oxide, which combines to a large extent the covering power of the white lead and the durability of the zinc oxide. The highest grades of mixed paints are made from white lead, zinc oxide, and pure linseed oil. Many substances are added to both lead and zinc paints as adulterants; such as chalk, barites, lead sulphate, kaolin, and a mixture of zinc sulphide and barites known as lithophone. High-grade paints should be free from such adulterants and from benzine or gasoline.

452. Colored Paints. These are made by adding some colored pigment to the mixture of white lead and zinc oxide. Some of the more common are yellow ocher and red ocher, which are oxides of iron and are comparatively inexpensive; chrome yellow, which is lead chromate; vermilion, a sulphide of mercury; Prussian blue; Paris green; chrome green, and so on. Shades of gray are produced by adding small quantities of lampblack to white paint.

453. Water Paints. These consist of a pigment suspended in water and held in place by some cementing substance such as glue or casein. These paints dry by evaporation. Most of the water paints are unsuited to outside work, although some of those made with casein have a fair degree of permanency when so used. Most of the kalsomines consist of chalk and some tinting pigment suspended in a thin solution of glue. Some of them, however, contain plaster of Paris, which forms a hard coating as it dries.

Whitewash is slaked lime mixed with water. When the mixture is spread on a surface the lime absorbs carbon dioxide from the air and forms calcium carbonate. Ordinary whitewash lasts only a short time when exposed to the weather. The preparation which is given below, known as Government whitewash, is fairly durable even for outdoor work.

454. Government Whitewash. Slake half a bushel of lime in warm water and cover it during the process to keep in the steam. Strain the liquid through a fine sieve or strainer. Add a peck of salt previously dissolved in warm water, three pounds of ground rice boiled to a thin paste and stirred in boiling hot, a half-pound of powdered Spanish whiting, and a pound of glue which has been previously dissolved over a slow fire. Then add five gallons of hot water, stir well, and let it stand for a few days protected from dust and dirt. It should be put on hot with small brushes. One pint of the mixture will cover a square yard.

455. Varnishes are used to provide protecting coats that are transparent and reveal the grain of the wood. As with mixed paints, there is a great variation in the quality of the varnishes on the market. A good varnish should stand water, and should not " dust" when scratched. Two general types of varnishes are oil varnishes and spirit varnishes.

The oil varnishes are made by melting a resin or gum and dissolving it in hot linseed oil, and then thinning the mixture to the proper consistency with turpentine. When made from pure linseed oil and a good gum such as opal, amber, or dammar, a varnish is produced that gives a tough, water-resisting film. Cheaper furniture varnishes, consisting of common rosin dissolved in oil, lack wearing quality.

Spirit varnishes are made by dissolving gums or resins in denatured alcohol or wood alcohol. These dry by simple evaporation of the solvent, and the gum is left unchanged except that it has been spread out in a thin film. The most common varnish of this kind is the well-known *shellac*. Spirit varnishes are not so durable as the oil varnishes.

456. Enamel Paints. These are made by grinding pigments in a good varnish instead of in linseed oil. The best white enamel paints consist of zinc oxide ground in a dammar varnish. Many of the enamel paints, however, are made from cheap varnishes and inferior pigments and fillers.

457. Black Varnishes. Varnishes used for coating iron are prepared by dissolving coal tar, pitch, or asphaltum in turpentine or benzine. They give excellent protecting surfaces where the black color is not objectionable. A very thin solution of these materials in benzine is often used as a stain for wood. The creosote stain often used on shingles is a coal-tar product, which preserves as well as stains the wood.

EXERCISES

Ex. 286. What is meant by the pigment in paints? By the vehicle? What is the best vehicle? How is linseed oil prepared? What happens to it when exposed to the air? What is meant by boiled oil? Why is the oil so prepared? What is meant by a drier and how does it act? Why is turpentine used in paints? What is Japan drier?

Ex. 287. Obtain a sample of linseed oil (from home if possible) and brush a little out into a very thin layer on a pane of glass. Examine this at the end of 24 hours and 48 hours. Does the oil harden completely? Is it suitable for paint making?

Ex. 288. What are the two best pigments for paint making? State advantages of each. Why is a mixture of both most desirable? How are colored paints made?

Ex. 289. Test a sample of paint from home for purity of the pigments as follows: (1) Remove all the oil from a tablespoonful of the solid sediment of the paint, by washing it several times by decantation with gasoline. Spread the residue out to dry. (2) Boil a portion of the dry material with strong acetic acid. A residue indicates barites or other adulterants. (3) Test another portion on charcoal for lead according to paragraph 252. (4) Test another portion for zinc according to paragraph 221. Record the results of your test.

Ex. 290. Of what do water paints consist? What is whitewash? Are water paints usually durable for outside work? Try the Government whitewash any place about the home where whitewash is used.

Ex. 291. How are varnishes made? What are the characteristics of a good varnish? What are spirit varnishes? Compare the finish from the two kinds of varnish. How are enamel paints made? What is the black varnish that is often used in coating iron?

CHAPTER XLVIII

CLEANING MATERIALS

458. THE operations of cleaning involve both physical and chemical processes. Dirt can sometimes be removed by means of brushing, shaking, or simple agitation under water. The dirt is first loosened by friction and then carried away by currents of air or water. It is usually held in place, however, by particles of grease or some sugary or gummy material and cannot be removed by strictly mechanical means. Sugary spots may be removed by warm water alone, but when the dirt is held by a film of grease, some substance must be used that will remove the grease. Soap is ordinarily employed to accomplish this end.

459. Soaps. It has been explained (303) that soaps are the sodium or potassium salts of the fatty acids, and are made by boiling fats or oils with the carbonates or hydroxides of sodium or potassium. Most of the soaps on the market are sodium soaps and belong to the class known as hard soaps. Soft soaps are made with potassium compounds but are not commonly found on the market. About the only familiar example of soft soap is the old-fashioned homemade soap which is manufactured from waste fats and lye.

The cleaning power of soap is to some extent due to the fact that it partially dissociates when dissolved in water, liberating a little free alkali which acts upon the grease to be removed. Probably its most important action is its power to emulsify the grease of the dirt spot. Dirt and small particles of the emulsified fat become thoroughly mixed with the suds, and upon rinsing are removed with the soap.

460. Pure Soap. Such a soap should contain no excess of alkali or of unchanged fat. A good test for free alkali is made by dropping a little phenolphthalein indicator on the freshly cut surface of a bar of soap. A coloration of the indicator shows the presence of free alkali. Unsaponified fat can be determined by drying a portion of the soap in the oven and then extracting the dry material with gasoline. If the gasoline leaves a greasy film upon evaporation, the presence of unchanged fat is indicated.

Toilet soaps especially should contain no free alkali, as the latter is injurious to the skin. Even in laundry soaps any large amount of free alkali is objectionable if the soaps are to be used for washing woolen or silk goods or any kind of colored material. Strong perfumes are so often used to cover the odor of the lye and other objectionable materials that it is advisable in general to avoid all highly scented soaps.

461. Hard Water Wastes Soap. When soap is used in hard water a scum appears on the surface of the water. This is due to the fact that the calcium and magnesium salts in the hard water react with the soap, forming calcium and magnesium salts of the fatty acids of the soap. These calcium and magnesium salts are insoluble in water and are so light that they float. As these salts are insoluble they have no emulsifying or cleaning power; hence the amount of soap that reacts with calcium and magnesium is wasted. Soap does not begin to lather until all the lime and magnesia have been removed from the water. In this case soap itself softens the water before it begins the real cleaning process; but it is an expensive method of softening water.

462. Foreign Ingredients in Soaps. Many of the soaps found on the market are adulterated. Washing soda, or sodium carbonate, in excessive quantities is present in many laundry soaps. The soaps especially recommended for hard water usually contain a large amount of sodium carbonate. While such a soap may not be objectionable for use on cotton goods in hard water, washing soda bought in this way is very expensive. Another common adulterant of laundry soap is sodium silicate, or water glass. This substance has some cleansing power, but its use results in an inferior soap. A few soaps contain a small percentage of borax.

Common rosin when boiled with an alkali forms a substance resembling soap. The yellow laundry soaps are usually made from a mixture of rosin and fats and are inferior to soaps made wholly from fats. Rosin soaps are not particularly objectionable for washing cotton goods, but with woolens they are likely to deposit rosin in the fiber and make the goods harsh to the touch.

Some soaps for use in cold water contain naphtha or kerosene. These substances are added for the specific purpose of assisting in dissolving the grease so that heat will not be necessary, and are not, therefore, to be considered as adulterants. Soaps are also frequently adulterated with substances of no cleaning value called fillers, which are added solely to increase the weight of the soap. Sodium sulphate, gypsum, whiting, chalk, and almost anything else that is cheap and bulky are used as fillers.

Water in excess of 25 per cent is also considered an adulterant. Too much water makes the soap soft so that it dissolves rapidly. A dry soap does not waste so readily; hence it is economy to buy laundry soap in quantities, unwrap the bars, and pile them loosely, so that the soap can dry as thoroughly as possible before being used.

463. Washing Soda and Other Alkalies. Where hard water must be employed, a judicious use of sodium carbonate results in a saving of soap. Just enough soda should be added to precipitate the lime and magnesia in the water. The soda should first be dissolved in a small quantity of water and this solution should be gradually added to the laundry water, care being taken to avoid any great excess. Caustic soda (soda lye) is sometimes used for the same purpose; but its use requires more care than ordinary washing soda. The use of either substance with silk or woolen goods is attended with risk, and colored goods of all kinds are likely to be injured by a slight excess of alkali in the wash water. Borax is occasionally used when a milder alkali than washing soda is needed, but it is much more expensive than soda.

Ammonia water is also used to soften hard water. It is the safest alkali to use in many cases because it evaporates very rapidly and does not remain in contact with the goods long enough to injure the fiber. The alkalies will destroy any materials made from oils or resins, hence strongly alkaline solutions should never be used on paints or varnishes.

464. Soap Powders and Scouring Soaps. Many of the washing powders on the market are nothing but washing soda. Others contain a small percentage of soap shavings mixed with a large amount of soda. It is probably more economical to buy the washing soda and soap separately than to purchase them in the form of washing powders.

Scouring soaps and powders consist for the most part of a small amount of soap mixed with washing soda and some scouring powder, such as whiting, chalk, powdered pumice, fine sand, or infusorial earth. The scouring material should not be so coarse as to scratch the article to be cleaned.

465. Solvents for Fats. When a grease spot occurs in a fabric which for any reason cannot be washed, it is necessary to use some substances that will dissolve the grease and carry it away. The most commonly used materials are gasoline and benzine. These liquids dissolve all the fats and are largely used in what is known as the dry-cleaning process. To get the best results some absorbing material must be placed beneath the fabric to be cleaned so that the grease can be washed through by the gasoline or other solvent. To add a little gasoline and rub the spot, merely spreads the grease into a larger and thinner film. Sometimes the goods are rinsed in several changes of gasoline to remove the grease.

FIG. 182. — Cleaning clothing with gasoline.

Ordinary ether is even better than gasoline for removing fats; but its high cost prevents its use except for certain special purposes. All the solvents ordinarily used for removing grease from clothing are highly inflammable, and their vapors make explosive mixtures with air. Great care should be exercised in using any of these materials, and they should never be used in a room in which there is a fire. Whenever possible the cleaning should be done out of doors.

A comparatively new substance known as carbon tetrachloride (CCl_4) is sometimes used to remove grease from fabrics. It has the advantage of not being inflammable.

Both gasoline and kerosene dissolve the calcium and magnesium soaps. Probably for this reason it is advantageous to place a little kerosene in the boiler in which clothing is boiled. For the same reason either substance is useful in cleaning bath tubs and basins where hard water is used.

466. Spots and Stains. These come from such a variety of causes that no remedy can be suggested unless the cause of the stain is known. At the best, however, it is practically impossible to remove stains from delicately colored fabrics without destroying the color of the dye as well.

Fruit juices and other acid substances frequently discolor dyed materials, especially the blues. The color can sometimes be restored by weak ammonia water if applied in time. Many fruit stains can be removed by hot water when fresh but are resistant to treatment when old.

Alcohol will remove grass stains (chlorophyll), and spots of varnish, and some paints. Consequently care should be exercised to avoid spilling alcohol, perfume, or other substance containing alcohol on any varnish finished surface.

Ink was formerly always made from iron tannate. This compound is soluble in weak acids, and can be removed, when not too old, by lemon juice and salt, oxalic acid, or even very dilute hydrochloric acid. The acid substance should be added cautiously and the material should be washed as soon as the color of the ink spot disappears. Many of the inks now on the market are made from aniline dyes and are not affected by the weak acids. For them there is no method of removal that will not affect the dye of the cloth as well. White goods can often be bleached to remove such ink spots.

EXERCISES

Ex. **292.** Review Exercises 176 and 177. What is the chemical composition of a soap? How do soaps clean?

Ex. **293.** Test several samples of both white and yellow soaps for free alkali by putting a drop of phenolphthalein indicator on the freshly cut surface of the soaps. Which show free alkali? Would they make good toilet soaps? Were any of the soaps especially recommended for hard water? Why should strongly perfumed soaps be avoided?

Ex. **294.** Do you use hard or soft water at home for washing? Explain how hard water wastes soap. What is the curd that floats on the top of hard water when soap is used in it?

Ex. **295.** Name some of the foreign ingredients sometimes found in soap. Why is it objectionable to have large quantities of sodium carbonate in soap? Why is rosin objectionable? What is meant by fillers in soap? What are naphtha soaps?

Ex. **296.** Dissolve a small piece of scouring soap in hot water. What is the nature of the residue? On a sample of any washing powder pour a little hydrochloric acid. Is sodium carbonate present?

Ex. **297.** Explain the softening action of sodium carbonate on hard water. Why should it be used cautiously with silks and wools? What advantage has ammonia water for use on woolens? Why should strong alkalies be avoided on paints and varnishes?

Ex. **298.** What kinds of spots can be removed by gasoline or ether? Upon what does their action depend? How are the best results obtained? What precautions are necessary in using gasoline or ether?

Ex. **299.** Make various spots on cotton or woolen goods, or use clothing already spotted, and attempt to remove the spots by the various methods mentioned in this chapter.

CHAPTER XLIX

INSECTICIDES AND FUNGICIDES

467. Losses due to Insects. It has been estimated that ten per cent of the value of farm crops is lost through the destruction caused by injurious insects, and that this loss amounts to four times that due to all destruction of property by fire. The farmer and the gardener, therefore, are interested in the methods of controlling injurious insects. These insects are divided into two general classes: those that eat the tissues of the plant, and those that suck out the plant juices and so injure the plant. Insects that bite may be killed by covering the plant with a so-called stomachic poison, which the insect swallows. The insects which merely suck the sap cannot be killed by the stomachic poison but are destroyed by contact poisons, which cause death by their caustic action on the body of the insect.

FIG. 183. — Spraying trees.

468. Stomachic Poisons. These poisons usually contain arsenic in some form. The oldest of these is Paris green (260), which is a compound of copper with arsenious acid and acetic acid. From one half to one pound of Paris green is used to fifty gallons of water. Paris green even in this dilute solution will sometimes scorch the leaves, and to prevent this action two pounds of slaked lime are added.

Arsenate of lead (251) is rapidly replacing Paris green as an insecticide. It is purchased either in the form of a paste or a powder and is used at the rate of from 2 to 4 pounds in 50 gallons of water. It is more adhesive than Paris green and does not injure the foliage.

Powdered hellebore is used to a limited extent by gardeners, especially for ornamental plants. The powder is mixed with hot water and then diluted and sprayed on the plants.

469. Contact Poisons. The first used of the contact poisons was kerosene emulsion. The kerosene is vigorously agitated with soapsuds until a jellylike emulsion is formed which can then be diluted with water without separating. It kills the insect by inclosing its body with a film of oil. If properly made, the emulsion is not injurious to the foliage.

470. Lime sulphur is probably used more than any other contact poison. Some of the insects, the San José scale for example, have the power of covering their bodies with a protecting scale and consequently a penetrating spray mixture is necessary to reach them. Lime-sulphur mixture seems to give the best results with these insects. It is made by boiling together lime and flowers of sulphur. The lime and sulphur unite to form the polysulphides of calcium, probably largely CaS_4 and CaS_5. When exposed to the air the polysulphides decompose, setting free sulphur. The insecticidal action of this spraying mixture is due partly to its caustic action on the insect, and partly to the poisonous effect of the nascent sulphur which is liberated. Lime-sulphur solution is now made commercially on a large scale and is not so generally manufactured on the farm as formerly. To kill the scale a strong solution is used when the trees are dormant. After the leaves appear, a weak solution is used for certain other insects and diseases.

471. Whale Oil Soap. This and other soaps made from fish oils are frequently used to destroy scale insects. The soap is dissolved in boiling water at the rate of two pounds to the gallon of water and is applied while hot. It is used principally for the winter treatment of scale-infected trees.

Common laundry soap dissolved in water at the rate of one bar of soap to two bucketfuls of water makes an effective contact insecticide for the small garden. A white soap free from rosin is to be preferred.

472. Tobacco is sometimes used for the destruction of insects on house plants. The stems or other refuse parts of the tobacco plant are steeped in water and the solution is sprayed on the plant. Sometimes in place of this decoction of tobacco a solution of nicotine sulphate is used. It will be remembered that tobacco contains an alkaloid known as nicotine (326). The salt formed by this alkaloid and sulphuric acid is the basis of such insecticides as Black Leaf-40. It is used especially to kill the aphis or plant louse.

473. Pyrethrum Powder. This powder, also known as *Buhach* and *Persian insect powder,* is composed of the pulverized flower heads of certain plants of the genus *pyrethrum.* It is a valuable insecticide for use in a small way when fresh, but soon loses its strength, especially if it is exposed to the air. It is used by dusting it on the plant or by steeping it in water and using the resulting liquid. The fumes of the burning powder are useful also in destroying insects in a confined space.

474. Gaseous Insecticides. Hydrocyanic acid (209) is sometimes used as an insecticide in greenhouses and in the orange groves of California. It is used in greenhouses in the following manner: the operator places dilute sulphuric acid in suitable vessels and then drops into each vessel a quantity of potassium cyanide wrapped in paper. He immediately steps outside and closes the house tightly. After a few hours the house is opened from the outside and thoroughly ventilated before any one ventures inside. In the orange groves each tree is covered with a tent for fumigation purposes. Hydrocyanic gas is a deadly poison and should be handled only by persons skilled in its use.

475. Carbon Bisulphide (110). This is used for killing weevils and other insects in stored grains. The liquid is placed in dishes at a number of places on the surface of the grain, about a teaspoonful being used for each cubic foot of space. The carbon bisulphide vaporizes, and the vapor, which is heavier than air, settles down through the grain. Carbon bisulphide is used also to exterminate rodents (110). It is inflammable and should not be used near a flame.

476. Sheep Dips and Fly Repellents. The materials used in dipping sheep for the purpose of killing their insect parasites are largely coal-tar products, cresol being the substance most commonly used. Light coal-tar oils and other coal-tar products are sometimes sprayed on cattle to repel flies. These repellents should be made of materials which, when they evaporate, leave no sticky or gummy substance on the hair of the cattle.

477. Insecticides for the Household. The common method of control for the common house fly is by means of some kind of fly paper. The sticky fly papers are coated with a mixture of common rosin and castor oil. The ordinary poison fly paper is impregnated with a dilute solution of sodium arsenite. Care should be exercised in using it, since children have been known to be poisoned by drinking the sweetened water in which the paper is placed.

A two per cent solution of formaldehyde (a teaspoonful of formalin to a cupful of water) is said to be as effective in killing flies as the poison paper and has the advantage of not being poisonous to human beings. A little honey or milk is added to the solution to attract the flies. If many flies are present, they may be killed by burning pyrethrum powder in the tightly closed room.

The favorite breeding places of flies are the garbage can, the manure pile, and other places where decomposing organic matter is found. Such materials should, when possible, be so handled as to prevent the flies from having access to them. It is said that hellebore, borax, or acid phosphate scattered on manure piles prevents the breeding of flies.

Sodium fluoride, NaF, is recommended for the extermination of cockroaches and red ants. It is scattered over the tables, sinks, or pantry shelves where these insects are found.

No insects cause the housewife more worry than the clothes moth which is so destructive to woolens and furs especially. The moth balls so commonly used consist of naphthalene ($C_{10}H_8$), a substance prepared from coal tar. It repels the moth and prevents it from laying the eggs on the clothing. If the eggs have already been deposited, the naphthalene has no value. The caterpillar which hatches from the eggs and does the real damage may be destroyed by dusting the garment heavily with pyrethrum powder. Sometimes the infected clothing is placed in a tight box and treated with carbon bisulphide as described for the treatment of grains (475). In the cities furs are sometimes placed in cold storage during the summer at a temperature so low that the eggs cannot hatch.

The common bedbug is destroyed by applying gasoline freely to every crack and crevice in which it is possible for the bug to hide. Repeated treatments of this kind will effectively control this pest. Fleas on dogs and other animals may be exterminated by the free use of pyrethrum powder. This powder is not poisonous to the higher animals, and its use, therefore, is attended with no danger.

Old houses occasionally become so overrun with insect pests that all ordinary methods fail. In such cases a resort to fumigation with hydrocyanic acid gas is justified. The same general method as that described for fumigating greenhouses is used. This gas destroys all animal life which is in the house, and is such a deadly poison that the greatest precaution should be observed in using it.

Fungicides. Plants are often injured by fungi that grow parasitically upon them. Some of these disease-producing fungi can be destroyed by spraying with a proper fungicide.

478. Bordeaux Mixture. Probably the best known of the fungicides is made by mixing a solution of copper sulphate with milk of lime (259) and diluting the mixture with water. It is the most common spray used on apple trees. As trees are attacked by insects at the time that they are suffering from fungus diseases, it is customary to add arsenate of lead or Paris green to the Bordeaux mixture so as to make the one mixture both an insecticide and fungicide.

479. Lime sulphur has marked fungicidal properties, and a dilute solution of lime sulphur is used in place of, or in addition to, the Bordeaux mixture for summer spraying.

Arsenate of lead is sometimes mixed with the lime-sulphur mixture.

480. Formaldehyde, or formalin, is commonly used to kill the smut fungus on oats and other grains. The commercial 40 per cent solution of formaldehyde is used at the rate of one pint to 20 gallons of water. Seeds are immersed in this solution for 10 minutes and then spread out to dry.

FIG. 184. — Treating seed potatoes with formalin to prevent scab.

A new method for treating oats for smut, which avoids the necessity of drying the seeds, has recently been reported by the New York State College of Agriculture. This method in brief is as follows: one pint of the 40 per cent solution of formaldehyde is diluted with one pint of water and placed in a quart hand sprayer. The oats are placed on a clean barn floor or on a tight wagon box. While the oats are being shoveled from one pile to another each shovelful is sprayed with the solution, giving one movement of the handle for each shovelful. After the oats are all treated in this way, they are ward it carried with it large quantities of rocks, grinding them against one another until they were reduced to particles of varying degrees of fineness. Such great force had this mass of ice with the rocks imbedded in it that it planed off the tops of the hills and carried the débris with it to the south. Later, when the climate became warmer, the ice melted, and this rock material remained behind to become a part of the soil. In some sections large masses of this material were left as long ridges called moraines.

FIG. 186. — The formation and transportation of soil by running water.

486. Action of Winds. In some parts of the world winds have played an important part in soil formation. Wind acts chiefly in transporting materials; but in localities where it blows with great violence it gathers up and sweeps along even coarse sand which, striking against rocks, slowly wears them away and is itself made finer. In the western part of this country may be seen instances of rock carving due to wind-driven sand (Fig. 187).

FIG. 187. — Peculiar rock forms carved by the action of wind and sand.

487. Chemical Weathering. Chemical changes in rocks take place at the same time that rocks undergo mechanical weathering, or pulverization. While the original rocks contain all the mineral elements required by the plant, these elements are present in unavailable forms. Such rocks as the granites are extremely insoluble. The rains that fall on these rocks contain carbon dioxide in solution, and the continued action of this weak acid results in a partial decomposition of the rock and a change of some of the material into soluble compounds. In this way some of the mineral matter finally becomes available as plant food.

488. Plants and Soil Formation. Plants play an important part in soil formation. The roots act both mechanically and chemically. Nearly every one has seen examples of the enormous force exerted by plant roots in breaking apart rocks (Fig. 188), when the plant gets started in a crevice or fissure. The same mechanical action is doubtless exerted by the roots on the rocks underlying some soils. The roots also secrete an acid substance, which acts on the rocks with which they come in contact; and when the plants die and the roots decay they leave in the soil numerous little channels which allow the passage of air and of water laden with carbonic acid.

FIG. 188. — **Growing roots of tree assisting in soil formation.**

489. Decaying organic matter produces humus, and this substance in its turn becomes an important factor in

PART III

SOILS AND FERTILIZERS

CHAPTER L

SOIL FORMATION

482. Soil. If soil is examined with the aid of a microscope, it is found to consist of particles of rock coated with a dark substance, which has been derived from the decomposition of organic matter. There is also present more or less vegetable matter consisting of fine roots and other parts of plants. The rock particles vary from stones of considerable size to grains of clay that are less than one five-thousandth of an inch in diameter. Every one is so familiar with the existing soil that it is hard to realize that nature required ages to form it and that numerous agencies contributed to the process of soil formation.

483. Weathering of Rocks. All soils have been formed by the decay or weathering of solid rock. One agency in the formation of the soil is change of temperature. A rock like granite, for instance, consists of different minerals cemented together. These minerals expand and contract at different rates when heated or cooled, and the result is that changes in temperature split the minerals apart. Indeed, even a mass of one kind of material is disintegrated by heat if the changes of temperature take place suddenly so that the surface and the interior of the mass are unevenly heated. The surface of a plate or of a piece of crockery, which has been used repeatedly in the oven, becomes covered with fine cracks caused by the uneven heating and cooling. The same thing happens in the case of some of the soil-forming rocks. Moreover, rocks are more or less porous and absorb water during the rains. In cold weather the water freezes and expands with great force and tends to break the rocks into pieces (6).


FIG. 185. — The disintegration of rocks by heat, cold, and frost.

484. Running water is an important agency in the weathering of rock. The water always carries some rock particles, and these rubbing upon the bed of the stream grind the stone to powder, which is carried away and deposited somewhere to form a soil. The more rapid the stream the greater its wearing effect.

485. Glaciers help to grind the rocks to powder. The action of the glaciers in Greenland, Alaska, and other countries, which may be studied at the present time, illustrates what happened in past geologic ages. Many thousand years ago a large area of the northern part of North America was covered with an immense glacier that pushed its way slowly down from northern Canada. As it moved south-ward it carried with it large quantities of rocks, grinding them against one another until they were reduced to particles of varying degrees of fineness. Such great force had this mass of ice with the rocks imbedded in it that it planed off the tops of the hills and carried the débris with it to the south. Later, when the climate became warmer, the ice melted, and this rock material remained behind to become a part of the soil. In some sections large masses of this material were left as long ridges called moraines.


FIG. 186. — The formation and transportation of soil by running water.

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FIG. 188. — **Growing roots of tree assisting in soil formation.**

489. Decaying organic matter produces humus, and this substance in its turn becomes an important factor in soil formation. The humus increases the power of the soil to retain water and to supply it to the plant; and since all the chemical changes by which plant food is made available take place more readily in the presence of sufficient moisture, it will be seen that this moisture-holding power of humus is a very important factor in soil fertility. During the decay of the organic matter, carbonic acid and other acid substances are produced, and these help to dissolve the mineral ingredients of the soil and change them into substances that can be absorbed by plants.

FIG. 189. — **Decaying organic matter assists in soil formation.**

490. Legumes in Soil Building. Plants belonging to the clover family, the so-called leguminous plants, through the bacteria that grow in the nodules or tubercles on their roots, are able to use the free nitrogen of the air as a source of food supply. When these plants die and become incorporated with the soil, the nitrogen which they have fixed becomes a part of the soil and is made available to succeeding plants. It has been estimated that from 50 to 150 pounds of nitrogen to the acre may be fixed in this way in a year. This power of the leguminous plants to accumulate nitrogen is probably nature's most important method of increasing the nitrogen content of the soil (168).

491. Animal Life in the Soil. In addition to the processes described above, the action of the earthworms and other forms of animal life found in the soil should be mentioned.

These organisms are supposed by some authorities to play a very important part in the working over of the soil and in its preparation for plant growth.

EXERCISES

Ex. **306.** Examine a sample of soil under a magnifying glass or a microscope. What material can you distinguish in the soil? What can you say about the sizes of the rock particles?

Ex. **307.** What is meant by the weathering of rock? Heat and cool a piece of granite repeatedly and note whether any change takes place. Heat a piece of granite in a flame and drop it into cold water. Does the sudden change of temperature affect the granite? Notice a plate at home which has been repeatedly heated in the oven. What causes the cracks on the surface? Explain how heat and cold help to pulverize the rock.

Ex. **308.** Recall Ex. 5. What can you say about the expansive force of freezing water? In what way does it help in soil formation?

Ex. **309.** Explain how running water helps in pulverizing the rocks. Do slow or rapid streams grind the rocks more readily? How do the streams form their valleys? What becomes of the materials which are ground to powder?

Ex. **310.** Explain the action of glaciers in forming soils. What part of North America was covered by the great glacier? What effect did the glacier have on the topography of the country? What is a moraine? Did the glacier cover your locality? If so, what evidence of that fact can you point out? Explain the action of winds in rock pulverization.

Ex. **311.** Explain how the insoluble silicates are made more soluble during soil formation. What is the original source of the nitrogen in soils? What can you say of the first plants which grew on the newly-formed soils? In what ways is organic matter important in soil formation? How do the roots assist in disintegration of the rocks? What is the chief function of the legumes in building soils?

Ex. **312.** How do earthworms affect the formation of soils? Can you find lichens growing on any of the rocks in your vicinity? What is the appearance of the rock under the lichen? How do you explain this appearance?

CHAPTER LI

KINDS OF SOILS

492. Physical Make-up of Soils. From the last chapter it will be seen that the completed soil consists of rock particles mixed with decaying vegetable matter, the decomposing remains of animals, and the various substances formed by chemical action from the rocks and organic matter. The rock particles are classified according to their size into gravel, sand, silt, and clay. Gravel is subdivided into coarse and fine gravel, and at least four grades of sand are recognized: coarse, medium, fine, and very fine. Silt also is divided into two classes: silt and fine silt.

The individual rock particles of clay are the smallest recognized by the soil physicist. They are so small that they cannot be distinguished by the naked eye, nor, indeed, can they be felt between the fingers. In other words the clay is entirely without grit, and when it is rubbed between the fingers only the smooth mass of clay can be felt, the individual particles being indistinguishable. As clays are very adhesive when moist, they adhere to tillage implements. Moreover, they absorb large amounts of water; and yet the individual particles lie so close together that water poured upon the surface of a clay often remains there a long time, soaking into it with extreme slowness. Silt is somewhat coarser than clay and the various grades of sand are still coarser. The gravels include the largest particles that are recognized as belonging to a true soil.

493. Classification of Soils. Soils are classified in two ways: (1) as to the method of their formation, and (2) as to their composition. According to formation, soils are divided into *sedentary soils,* or those that were formed where they now exist; and *transported soils.* Sedentary soils are subdivided into *residual* and *cumulose* soils. Transported soils include *alluvial, colluvial, drift,* and *æolian* soils. According to composition, soils are classified into *clay, sand, loam,* and *peat* soils.

494. Residual soils, which have been made from the decay of the rocks on which they lie (Fig. 190), partake more or less of the composition of the underlying rocks. They have usually lost considerable of their soluble constituents through the solvent action of rains. These soils are not generally very deep, the underlying rock being comparatively close to the surface. Residual soils may be fertile or not, their degree of fertility depending upon the kind of rock from which they were formed.

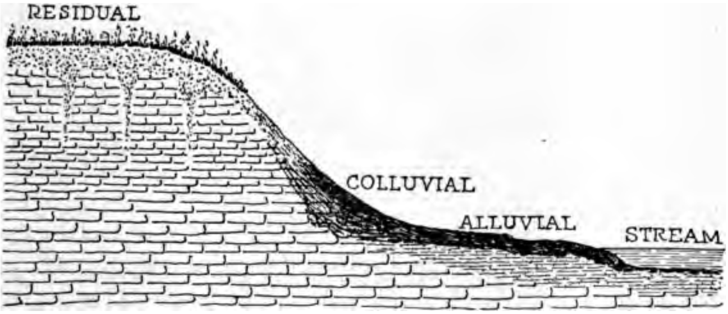

FIG. 190. — The formation of residual, colluvial, and alluvial soils.

495. Cumulose soils are those formed in swamps and marshes. They consist largely of organic matter which has come from the partial decay of the marsh plants. They contain also the earth which has been washed down from the surrounding higher lands. Muck and peat are examples of cumulose soils.

FIG. 191. — The formation of cumulose soil.

496. Alluvial soils are those that have been carried by water and deposited some distance from their original source. They commonly show more or less distinct layers as a result of the fact that the coarser particles naturally settle first, while the finest particles are the last to be deposited. The soils in river valleys are alluvial and have been carried down by the stream during the flood season and deposited as the velocity of the current decreased. Alluvial soils are usually fertile, but it will be seen that the character of these soils varies with the character of the rock material of the uplands from which they are derived.

497. Colluvial soils are formed on the lower slopes of hillsides. They are composed of particles of various sizes that have moved down the hillside under the force of gravity. (See Fig. 190.)

498. Drift soils are those that have been transported by glaciers. A large part of the northern United States is covered by drift soils that were carried down from the north by the great glaciers which at one time covered this region. Drift soils are characterized by the presence of bowlders and rounded pebbles. They vary considerably in character, and many different kinds of soil may be found in a single farm located in the glaciated region. Drift soils are usually very productive.

499. Æolian soils, or loess, are those that are composed of particles transported by the wind, and are, therefore, sometimes called wind-formed soils. It is supposed that considerable areas of soils in the central United States are wind-formed. They vary in thickness from a few feet to over 100 feet, and are of considerable agricultural value.

500. A clay soil is one that contains over sixty per cent of clay particles. It is the hardest soil to work, since it is sticky when wet and so hard that it can hardly be pulverized when dry. In very dry weather clay soils crack and form openings that allow excessive evaporation of water, a condition which dries and injures the roots, sometimes breaking them. Clay soils, unless well drained, are likely to be cold and unresponsive. These soils are usually high in potential plant food, especially potash, but require careful handling to enable the plant to make use of this food material. They are usually retentive of added plant food and are, therefore, soils that can be liberally fertilized without fear of great loss of the applied material.

501. Sandy soils are those containing a very large proportion of sands (75 per cent or more). They are just the opposite of clay soils, being too open and porous, while the clays are too compact and impervious. They hold but little water, and crops growing on them are likely to suffer in hot, dry weather. These soils are usually low in fertility and have little power to retain added plant food, since the soluble material in manures and fertilizers used on them is likely to leach through them. No soil is so poor that it cannot be made to grow a crop, and even sandy soils can be made productive by the liberal use of organic matter, and the addition, if necessary, of lime, phosphoric acid, and potash. They are warm, easily worked soils, and if properly handled are often profitably used, especially for early truck crops.

502. Peat and muck soils contain very large amounts of organic matter, some of them having as much as 80 per cent of such matter. They are found in the beds of former lakes or swamps, and since they are formed by the partial decay of vegetable matter under water, they usually contain but little earth. Peat is an intermediate product between vegetable matter and coal, and perhaps in course of time it would be converted into coal. The name muck is sometimes applied to a soil in which the organic matter is in a more advanced state of decay than it is in true peat. A muck contains more earth than the peat does and is more compact. These soils are high in potential nitrogen, but are usually exceedingly low in potash. When they are well drained and are fertilized with phosphorus and potash they are generally fertile soils. Many of the black onion soils belong to this type.

FIG. 192. — Blocks of peat being dried for use as fuel.

503. Loam is a soil consisting of a mixture of clay, silt, sand, and organic matter and for most purposes is the most desirable type of soil. Loam soils are usually well balanced, since they hold moisture well, are well supplied with plant food, and have considerable ability to retain such soluble plant food as may be added to them. They allow the air to circulate through them more freely than clay soils do, but they are not so objectionably open as are sandy soils. They are easily worked and have comparatively little tendency to bake or crust on the surface. They are well suited to most crops and respond well to fertilization. Several subtypes are recognized: (1) heavy clay loam, (2) clay loam, (3) silt loam, (4) sandy loam, (5) light sandy loam. Taken in the order given above they contain from first to last decreasing quantities of clay, and increasing quantities of sand, the heavy clay loam having the most clay and the light sandy loam the most sand. These subtypes naturally partake of the characteristics of their components, so that the heavy clay loam shows in a large measure the properties of a clay, and the light sandy loams are only a step removed from the sandy soils. The others are intermediate between the two extremes. The loams as a whole represent the more common types of farm soils.

504. Light and heavy soils are names which do not, as might be supposed, refer to the actual weight of soils, but to the ease with which they may be worked with tillage implements. A sandy soil, for instance, is called a light soil, although it actually weighs more per cubic foot than any other type of soil.

EXERCISES

Ex. 313. What are the four general classes of rock particles in soils? How are these classes subdivided? Which particles are the smallest? Work up a little pure clay with water and rub it between the fingers. Do you feel any grit? Do the same with a sandy soil, and note the difference. Place about two tablespoonfuls of fine soil in a quart fruit jar three fourths full of water. Shake the jar vigorously for several minutes, let it stand for one minute, and pour the muddy water into a second jar. The sediment remaining in the first jar is composed almost entirely of sand. Examine it carefully. When the second jar has been standing five minutes pour off the muddy water into a third jar; add more water, shake this jar vigorously, and after it has stood five minutes pour off the water. The sediment remaining in the second jar is largely silt. Let the third jar stand at least two hours and then pour off the water. The residue in this case is largely clay. Note how fine the particles are. Examine them under a microscope. Can you find samples of clay, silt, and gravel in your vicinity?

Ex. **314**. What are the two general classes of soils according to formation? Are the soils in your vicinity sedentary or transported or both? What are residual soils? Cumulose soils? Can you find either type near the school? What is meant by alluvial, colluvial, drift, and æolian soils? How many of these can you find near your home? How can you tell a drift soil? Walk along a stream and note how the rock particles of varying sizes are deposited by the stream as the current becomes less rapid.


FIG. 193. — A soil auger.

Ex. **315**. What are the chief characteristics of a clay soil? Of a sandy soil? Of a muck soil? Why are clay soils more difficult to cultivate than sandy soils? Which hold more water? Which contain more plant food? Are there any areas of muck soils in your locality? For what are they used?

Ex. **316**. What is meant by a loam? Why are loam soils desirable? What general classes of loam soils are recognized? How many of them can you find in your locality? Bring samples of as many different kinds of soil as you can find to the school and classify them according to this chapter. (The best method of obtaining a sample of soil is by means of the soil auger (Fig. 193), which is made by welding a three-eighths inch gas pipe to a one and one-fourth inch wood auger. The class should take an excursion over the neighboring farms and by the means of the soil auger locate as many soil types as possible.)

CHAPTER LII

RELATION OF THE SOIL TO PLANTS

505. Permeability to Plant Roots. The soil furnishes an anchorage for the plant roots and enables them to hold the plant in an upright position. Since the roots of the ordinary farm crops must penetrate into the soil to a distance of from two to ten feet in order to obtain the necessary amount of plant food, the permeability of the soil is a matter of great importance. A soil that is so compact as to hinder the growth of the roots, or one that has near the surface a hardpan through which the roots cannot penetrate, seriously interferes with the full development of the plant. Since the cells of the growing points of the roots need oxygen for respiration, it follows that the soil must be permeable to air as well as to the plant roots.

506. Water-holding Capacity. Attention has already been called to the enormous amount of water required for plant growth (346). The plant needs water during the entire growing season, and as the rains are irregular and often come weeks apart, the soil must act as a reservoir for moisture. The water-holding capacity of the soil, then, is a factor of great importance. The proper condition of moisture in the soil is the most important single factor in determining the fertility of the land, and the failure of soils to produce good crops is more often due to lack of available water than to any other one cause. Any plan that will increase the capacity of the soil to store water is desirable.

507. The Soil Supplies Nitrogen. All plants except the legumes are dependent upon the nitrates in the soil for their supply of nitrogen. Although a good soil may contain from three to five thousand pounds of nitrogen to the acre in the surface foot, only a few pounds of it is in the form of nitrates. Nearly all the nitrogen is present in organic matter; but in these complex organic compounds it cannot be utilized by plants. Nitrification (168), therefore, is necessary to the maintenance of the fertility of the soil. Nitrification takes place only when the temperature of the soil is more than five degrees above freezing, and becomes more rapid with rise of temperature. Hence, it ceases during the winter months and is most vigorous during the hot months of midsummer. The nitrifying bacteria cannot live without a sufficient supply of oxygen, and for this reason stirring the soil to introduce air increases the rate of nitrification. Moreover, these bacteria cannot thrive in a soil that is acid, hence carbonate of lime is essential to nitrification. So vital is the process of nitrification to the growing crop that successful agriculture depends largely upon providing proper conditions for rapid nitrification.

508. Denitrification. While the nitrifying bacteria may be said to be the farmer's friends, there are, unfortunately, other organisms in the soil that produce evil results. One class of these, known as denitrifying bacteria, decomposes the nitrates, and, perhaps, some other nitrogenous compounds, with the final result that the nitrogen is set free and returned to the air in its elemental condition. This process, of course, robs the soil of a part of its nitrogen, and is especially unfortunate because it removes the part that was most readily available to the crop. Denitrification can be prevented by providing the conditions that favor rapid nitrification.

509. Fixation of Nitrogen. Leguminous plants are not absolutely dependent upon the nitrates of the soil. These plants use the nitrates as long as they are available; but when the soil fails to supply nitrates in sufficient quantity, they depend upon the nodule-forming bacteria, and thus indirectly make use of free nitrogen. The fixation of nitrogen is not a function of the legume itself, but of the bacteria that produce the nodules; and in the absence of these organisms the legumes are quite as dependent upon the supply of nitrates as are the other families of plants. In soils very rich in nitrogen the root tubercles may not be formed on legumes even when the proper bacteria are present. In ordinary soils a crop of clover obtains one third of the nitrogen from the soil and two thirds from the air.

FIG. 194. — **Effect of inoculation of soil on growth of soy bean.** The plant on the left was grown in inoculated soil; that on the right, without inoculation.

510. Inoculation of Soils. Experience has shown that not all soils contain the bacteria necessary to the fixation of free nitrogen by legumes. These bacteria may be introduced into a field by scattering on it a small quantity of soil from a field in which the same legume has been successfully grown, and then harrowing it in. Nitrogen-fixing bacteria that grow on one kind of legume will not thrive on all other legumes. Therefore the soil from a red clover field is not suitable for inoculating soil for soy beans or alfalfa. Soils on which sweet clover grows, however, may be used to inoculate for alfalfa.

Inoculation of the soil is of undoubted use in some cases; but there is danger of overestimating its value. It must not be regarded as a panacea for all the ills of the soil. Inoculating a soil simply introduces the nodule-forming bacteria, and if the failure of a leguminous crop was due only to absence of these bacteria, inoculation will be beneficial. It will in no wise overcome failure resulting from bad seed, improper preparation of the ground, adverse weather conditions, or acidity of the soil; and the farmer should assure himself that the soil conditions are as favorable as possible before he attempts inoculation.

511. Mineral Elements. It has been shown that seven elements are necessary to plant growth (350). Of these, only phosphorus, potassium, and calcium need be considered, since experience seems to show that the others are present in all soils in sufficient quantities for maximum crop production. The amounts of nitrogen, phosphorus, and potassium removed by crops seem insignificant when compared with the total quantities of these elements that are present in a good soil. The grain and straw of a thirty-bushel crop of wheat, for example, remove from an acre 60 pounds of nitrogen, 10 pounds of phosphorus, and 35 pounds of potassium. The first foot of a good loam soil, however, contains to each acre about 5000 pounds of nitrogen, 2200 pounds of phosphorus, and 35,000 pounds of potassium. Such a soil, therefore, has sufficient nitrogen for 80 crops of wheat of 30 bushels each, phosphorus enough for 220 such crops, and potassium enough for 1000 crops. Yet it is well known that such a soil would not produce even eighty such crops of wheat in succession.

512. Chemical Analyses. A chemical analysis gives the total amounts of nitrogen, phosphoric acid, and potash in the soil, but does not indicate what part of these substances is available to the plant. The greater proportion of these substances is locked up in insoluble compounds, in which form the plant is incapable of using them. Smaller quantities have been changed by the forces of nature into a condition in which they are available to plants. While the amounts of these materials removed by the crop seem insignificant when compared with the total plant food in the soil, they may be very large in comparison with the available part. The unavailable, or potential, plant food is gradually being made available, but not with sufficient rapidity to replace that removed from the field at harvest. It will thus be seen that the *present fertility* of the soil depends not upon the potential plant food it contains but upon that which is immediately available to the plant. A chemical analysis is of value, however, in showing the potential possibilities of a soil.

513. Limiting Factors. Since a definite amount of each element of plant food is required for a certain yield, and since none of the elements can be replaced by another (350), it follows that the crop produced will be limited by the quantity of the essential element present in least proportion, compared with the requirements of the crop. In other words, if a field of corn can obtain phosphorus sufficient for only half a crop, no more than this can be produced no matter how much of the other elements of plant food is present. The substance which thus limits the crop production is said to be the *limiting factor* for that soil. The maximum yield of a particular field is often determined by some one factor, and it is important to discover what that limiting factor is and to apply a remedy. The lack of sufficient water is the most common limiting factor; phosphorus probably ranks second; and either calcium carbonate or nitrogen ranks third. Potassium is the usual limiting factor in peat and muck soils.

514. Nature's Methods Contrasted with Those of Man. The amount of vegetation which the soil can produce has been constantly increasing. Under natural conditions this growth is not removed from the ground, and the plant food therein is again made available, so that the land is constantly increasing in fertility. Thus the fertility of virgin soils is the result of accumulations due to a variety of forces acting through countless ages, during which little has been removed from the soil while much has been added thereto.

Man, on the contrary, has reversed this process, and while adding little to the soil has removed much therefrom. Through the constant harvesting of crops, and the leaving of the ground bare and exposed to the action of the elements, he is rapidly depleting nature's store of food, and the yield steadily becomes smaller. To prevent exhaustion of the soil it becomes necessary, therefore, for him to assist nature in making potential plant food available and to return to the soil at least a part of that which he removes in the crops.

EXERCISES

Ex. 317. Why must the soil be permeable to plant roots? What does the plant obtain from the soil? Must the oxygen of the air penetrate into the soil? Examine a section of soil in an excavation.

Ex. 318. Place 100 grams of a thoroughly dried soil in a beaker, add 100 cubic centimeters of water, and stir for several minutes. Place a folded filter paper in a funnel and moisten the filter. When water no longer drips from the filter, place a 100 cc. graduated cylinder under the funnel and pour the mixture of soil and water into the filter. When water ceases to drip from the funnel measure the water which has run through. The number of cubic centimeters of water in the cylinder subtracted from 100 represents in percentage the water-holding capacity of the soil. Repeat the exercise with a sandy, clay, and loam soil. Compare results. Why is the water-holding capacity of the soil important?

Ex. **319.** In what form do most plants use nitrogen? Is the process of nitrification very important? At what time of the year is nitrification most active? What are the soil conditions necessary for nitrification? Why is denitrification undesirable? Under what soil conditions does denitrification proceed most rapidly?

Ex. **320.** Dig up a clover plant or other legume and wash the dirt from the roots. (Do not shake dirt loose.) What connection have the nodules on the roots with the nitrogen supply? Will legumes form nodules in soils rich in nitrogen? Is the fixation of nitrogen a function of the legume or of the nodule bacteria?

Ex. **321.** What proportion of its nitrogen does clover obtain from the air and what proportion from the soil? Make a collection for the school of the different root nodules in your neighborhood and preserve them in water with a 10 per cent solution of commercial formaldehyde.

Ex. **322.** Do all soils contain the proper bacteria for the fixation of nitrogen by legumes? Do the same kind of bacteria work with all legumes? How may the bacteria be introduced into a soil in which they are lacking? To what extent is inoculation of soils valuable? Will inoculation take the place of other preparation of the ground?

Ex. **323.** Which of the elements of plant food are likely to be deficient in the soil? What does a chemical analysis tell about the fertility of the soil? Explain what is meant by available and potential plant food? What is meant by a limiting factor in plant growth? What are the more common limiting factors in plant growth? Contrast nature's methods in regard to the soil with those of man.

CHAPTER LIII

SOIL WATER

515. Structure of Soil. A careful examination of the soil shows that it is made up of grains of various sizes fitted somewhat loosely together, with spaces between them which are as variable in size as the grains themselves. In a loam soil in good tilth the spaces between the grains represent about half the total volume of the soil.

516. Ground Water. In the surface soil the spaces between the grains are usually filled with air, but at some distance below the surface the spaces are entirely filled with water. This water is known as *ground water*. The upper surface of the ground water is called the *water table*. The exact height of the water table may be ascertained by sinking a hole to such a depth that water will stand in it, the level of the water in the hole being practically that of the water table. It is ground water that supplies wells and springs.

517. Free Water and Film Water. After a rainy period the ground may become saturated with water. In such a case the water table is at the surface. The water table gradually sinks, and after a few days of dry weather it may be a foot or more below the surface. The soil water that has drained downward by the pull of gravity is called *free,* or *gravitational, water*. The water that remains above the water table is called *film water*. The following experiment will make clear the difference between free water and film water. Cork tightly the opening at the bottom of a flower pot. Add soil to the pot until it is nearly full. Then pour in water until the soil is saturated. When this condition is reached the water table is at the surface of the soil. Now remove the cork, and much of the soil water will drain off. This is free water. The water that remains in the soil exists as thin films of moisture around the grains of soil. These films join where the soil grains touch one another and really make one continuous film of moisture throughout the soil. This film moisture is also called *capillary water* because it moves through the soil by capillary attraction as oil moves upward through a lamp wick.

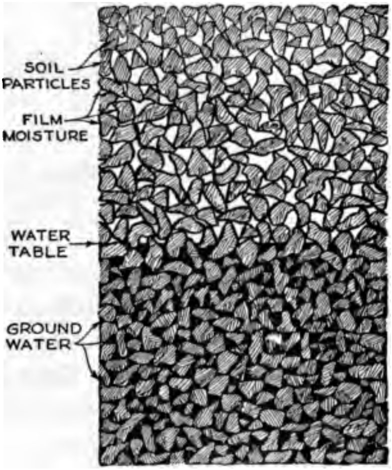

FIG. 195. — **Sectional view of soil showing ground water and film moisture.**

518. Movement of Film Water. The tendency in the soil is to maintain an even thickness of film water over all the soil grains; consequently if anything happens to decrease the thickness of the film in one part of the soil, water will move toward that point to restore equilibrium. When water evaporates from the surface of the soil, more water moves upward to replace the loss. This movement is known as the capillary rise of water. It will thus be seen that evaporation not merely dries the surface of the soil but actually pumps up water from the lower layers and affects the water content of the soil for some distance below the surface. The rate and the amount of capillary movement depend upon the structure of the soil. The coarser the soil the more rapid the rise of water by capillarity, but the finer the soil the higher the water will be lifted. A very coarse sand may raise the water only a few inches, a fine sandy loam will lift it a few feet, while in a clay soil the water will rise at least twenty-five feet. If the soil is made more compact, the capillary rise of water is increased.

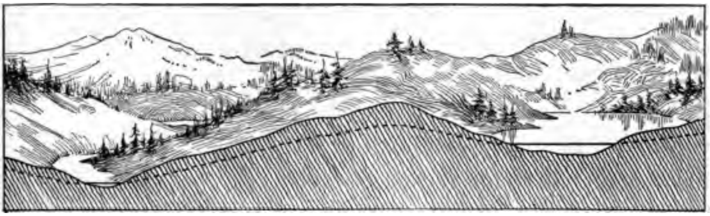

FIG. 196. — **Relation of the water table to contour of the land.**

519. How Plants Get Their Water. The roots of the plant push their way down between the soil grains, branching more or less and spreading throughout the soil. The root hairs at the growing points of the roots, which are the absorbing organs of the plant, work their way in between and around the small soil grains, adhering very closely to them. The root hairs absorb the water of the film moisture, and as they reduce the thickness of the film, more water diffuses to the point of absorption. Since the root hair gets all its water and food from the film on the surface of the soil grains, film moisture is a desirable form in which to have water in the soil.

520. High Water Table Objectionable. The crops are sure to suffer when the level of the ground water is near the surface of the soil (Fig. 197). A high water table limits the feeding space available to the plant and, consequently, the amount of food it can obtain. The plants that are of agricultural importance must have their roots supplied with air, and such plants do not send their roots below the water table because the spaces between the soil grains below this level are filled with water, a condition which prevents the entrance of air. The depth to which the roots will go, then, depends upon the position of the water table.

Free water near the surface is also objectionable because it makes the soil cold. It requires much more heat to warm water a certain number of degrees than to raise the temperature of an equal weight of the dry matter of the soil the same number of degrees. Hence a soil that contains much water is harder to heat than one that is comparatively dry.

A very wet soil causes plant food to become locked up in unavailable forms, and in some cases brings about the production of compounds that are actually poisonous to the desirable plants. An excessive amount of water in the soil also dilutes the plant food in solution and makes it more difficult for the plant to procure sufficient nourishment.

One of the most important considerations in this connection is the fact that the presence of free water in the soil prevents nitrification and promotes denitrification. In water-logged soil, nitrates are rapidly decomposed, the nitrogen being given off to the air in the free, or elemental, condition; and for this reason not only is nitrogenous food in the soil destroyed, but the application of nitrogen fertilizers to such a soil results in great waste of this valuable element.

521. Underdrainage. When the water table, during much of or all the growing season, is nearer than three feet to the surface of the ground, some system of underdrainage becomes necessary, if the best results in crop production are to be achieved. This is best accomplished by means of drain tile. The benefits of tile drainage are summarized in the following paragraphs.

The water table is lowered to the level of the drain tile, the water running off through the tile instead of remaining in the soil as stagnant water. The plant roots can now penetrate the soil to a greater depth, since air follows the water as it percolates downward, thus ventilating the soil. This ventilation is of great importance and it is continuous, since each rain forces some of the old air out of the soil, and the new air following the rain takes its place.

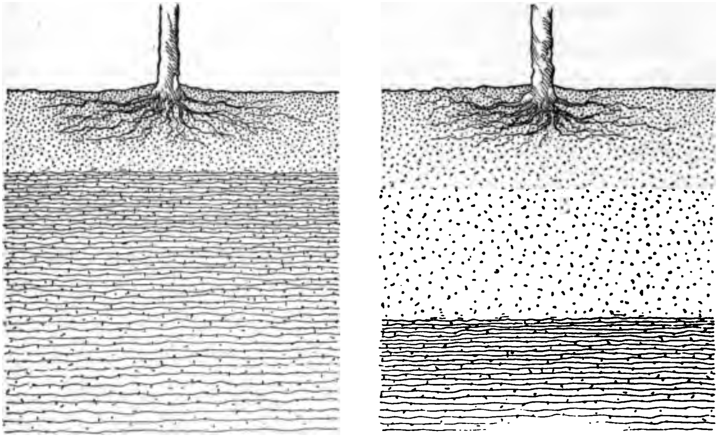

FIG. 197. — **Effect of lack of drainage on root growth.** The high water table shown on the left limits the feeding space of the roots, and in time of drought, as shown on the right, the plant suffers because of its shallow roots and lack of moisture.

In a well-drained area, more of the rain will soak into the soil instead of running off the surface; thus surface washing of the soil will be prevented to some extent. Ideal conditions demand that there shall be no run-off at all, and the farmer should strive to attain as nearly to this ideal as is practicable.

In some countries, where terrace farming is followed, the run-off is almost entirely eliminated.

The spring rains are usually warm, and it is desirable to have them percolate through the soil. Tile-drained soils warm up earlier in the spring, stay warm later in the fall, and maintain a higher temperature throughout the growing season than do undrained soils. This is partly due to the fact that more evaporation takes place from the surface of the undrained soil (8). For these reasons crops on well-drained soils have a longer season for growth.

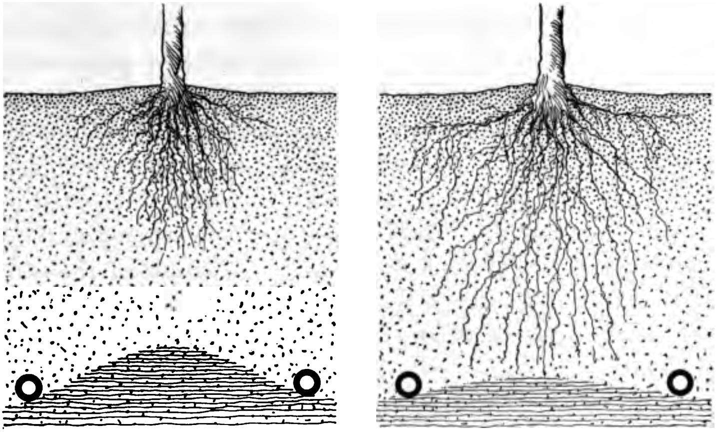

FIG. 198. — Effect of proper drainage on root growth.

Tile drainage promotes nitrification and prevents denitrification. The decay of organic matter, resulting in the production of nitric acid, is an oxidizing process and can take place only in a well-ventilated soil.

Many thousands of square miles of waste land have been reclaimed by means of tile drainage. In the case of swamps, marshes, and ponds the water table is actually at or above the surface of the ground, and such lands will not produce the ordinary farm crops at all until they are drained.

522. Drainage Reduces Injury from Drought. Paradoxical as it may seem, underdraining increases the amount of water available to the plant. The crop depends almost entirely on the capillary or film moisture for its supply of water, and the roots do not enter that part of the soil containing free water. Lowering the water table greatly increases the total amount of film moisture, as all that part of the soil from which the free water has been removed is capable of holding capillary water. Thus it will be seen that while the total amount of water in the soil is decreased by drainage, that amount which is of use to the plant is made much greater. Drainage prevents injury from drought also, by allowing plants to make deeper root growth, hence they are not so easily affected by the extreme drying of the surface of the ground that takes place in times of scanty rainfall.

523. No Useful Water Lost Through Tiles. It will readily be seen that tile draining determines the highest point the water table can reach, but in dry weather the level of the ground water may be much below the drain. It is sometimes feared that for this reason a part of the water from summer rains may be lost through the tile. Experience has shown, however, that the water does not percolate into the drain, as some suppose, but that the drains remove water only when there is sufficient rain to raise the water table to the level of the tile. It is simply the excess of water that is removed by the underdraining, and not the part that is of importance to the plant.

524. Drainage Sometimes Beneficial on High Lands. Strangely enough, experience has shown that it is not merely low-lying soils that are benefited by underdraining. In many cases heavy clay soils in elevated positions, especially if underlaid by rather impervious subsoils, are greatly improved by tiling. In such soils the percolation is so slow that practically the same effect is produced as would be expected if the general level of the ground water were near the surface. These soils are made more mellow by drainage and respond more readily to early tillage. In fact, it may be said that it will pay to tile drain any soil that has a clay subsoil, no matter what the elevation or slope of the land.

525. Draining the Farm. The laying out of a complete system of drainage calls for the services of a skilled drainage engineer, especially if the farm is very level. The more fall there is to the land, the easier it is for an unskilled person to lay drains that will work. Since the tile must be so laid

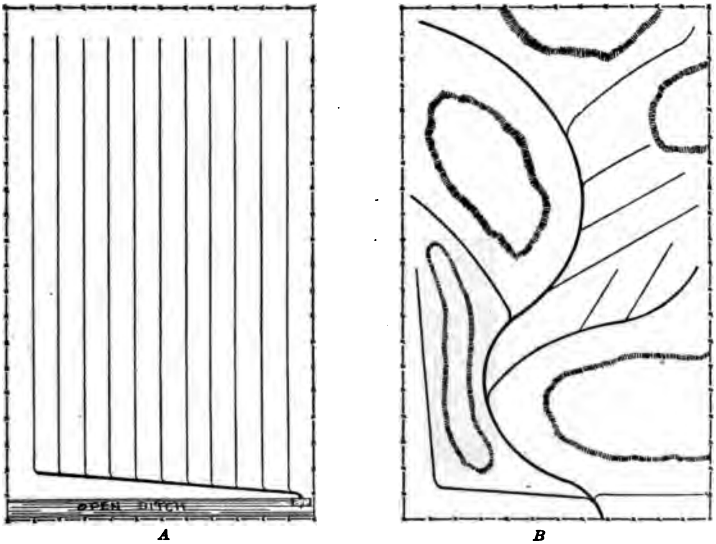

Fig. 199. — *A*, Drainage system for level land. *B*, Drainage system where there is much variation in the slope of the land.

as to give a uniform fall from the highest point to the outlet, the planning of the system should be preceded by a survey of the farm for that purpose. When the field to be drained is broad and level (Fig. 199 *A*), the main drains and laterals are uniformly spaced; but often the ground has natural drainage lines (Fig. 199 *B*) along which it is usually best to run the ditches, especially the main lines.

The size of the tile should vary with the area to be drained and with the grade at which the tile is laid. With an average fall of one inch to one hundred feet, a four-inch tile will drain ten acres. With a greater fall this size of tile would remove the water from a larger area.

FIG. 200. — Drainage machine at work.

The depth at which the tile is placed and the distance between the lines of tile depend upon the character of the soil. The more compact the soil the nearer together the lines of tile and the less the depth. In very heavy clays the drains should not be farther than two rods apart, while in lighter soils the distance apart may be twice as great. Clay soils should have the drains placed about thirty inches deep; but in light soils the tile may be as much as four feet below the surface.

Unglazed clay tile are the more common. Glazed tile may be used, however, since all the water enters at the joints and there is no advantage in having porous tile. Machines are now on the market for making tile from cement.

The ditches for laying the tile are often dug by hand, and various special tools have been devised for use in ditching as well as for laying the tile. Machines worked by horse power and by tractors (Fig. 200) are used in ditching, when the character of the soil permits.

FIG. 201. — **Laying tile in ditch.**

526. Irrigation. There are sections of this country and of other countries that receive no rainfall or such a limited amount that crop production is impossible without the aid of some system of irrigation. The last few decades have seen thousands of square miles of the earth's surface that were formerly barren made productive by means of irrigation. Water for this purpose is sometimes obtained from wells and springs, but most of it comes from running streams. In sections where the streams become too low during the summer months the water is collected in reservoirs from which it is drawn as needed. In some cases where there is a plentiful supply of water it is merely diverted from the river into canals from which the fields are supplied.

Irrigation farming has one advantage over farming in humid climates, namely, that the farmer has the water supply for the crops directly under his control. Water can be added at the time it is needed by the plants and in the right quantity, without dependence upon the uncertainties of natural rainfall. Under favorable conditions the expense of providing the water is more than repaid in the larger crop produced.

527. Methods of Irrigation. Two general methods of distributing the water over the field are used in the arid sections of this country. In the first method the water is flooded over the ground in as even a layer as possible. For best results with this system the ground must be quite level. This method is used on such crops as alfalfa or other hay crops. On cultivated ground it is not so desirable.

FIG. 202. — Irrigation by furrows.

The second method is more commonly used, and consists in distributing the water by means of furrows (Fig. 202). These furrows are opened at distances varying with the character of the soil. They have sufficient fall to cause the water to run slowly through them and so soak into the ground as it passes along the furrow. This method seems to be the most economical of water and labor for use on cultivated lands. The western fruit orchards are irrigated in this way.

528. Irrigated Soils Must be Underdrained. Even in arid climates, soils which are irrigated must be tile drained after a few years of irrigation. One reason for this is that the wonderful yields produced by the application of water have in some instances led to overdoing irrigation, with the result that the soil becomes water-logged. Many areas on the delta of the Nile, for example, have become water-logged since the introduction of perennial irrigation, and many of these tracts are now being underdrained.

FIG. 203. — Irrigation by sprinkling.

Another reason why drainage is necessary in arid sections is that irrigation without drainage brings about the accumulation of an injurious quantity of alkali salts near the surface of the soil. The water dissolves the salts and brings them to the surface by capillarity; and then as the water evaporates the alkalies become concentrated near the surface. By putting in tiles and using plenty of water the excess of alkalies can be washed out of the soil.

529. Irrigation in Humid Climates. Although there is no doubt about the value of irrigation in arid climates, its usefulness in that part of this country that lies east of the Mississippi River is not so generally recognized. Eastern market gardeners find, however, that it pays in most seasons to do some irrigating. Some of them use a plan much like the furrow system, but many have adopted a sprinkling system similar to the one shown in Fig. 203. This method of applying water appears to be very successful for market gardens in the central and eastern states, where it merely serves to supplement a naturally good rainfall. Moreover, the gains in yield through irrigation of certain field crops at the Wisconsin Experiment Station indicate that even in the Middle West it might pay to irrigate such crops on farms where water could be obtained at little cost.

EXERCISES

Ex. **324.** Fill two glass tumblers to within one half inch of the top, one with a clay soil and the other with a sandy soil. Compact the soil by gently tapping the tumblers on the desk. Pour water on slowly from a graduated cylinder until it stands just at the level of the top of the soil. Note the difference in the amount of water required by the two soils. The water absorbed represents the pore space of the soils. Which has the more pore space, a coarse-grained or a fine-grained soil? How much pore space does a good loam soil contain?

Ex. **325.** Perform the experiment described in 517. What is meant by film moisture in soils? By ground water? By the water table? At what depth is the water table where you live? How can you tell?

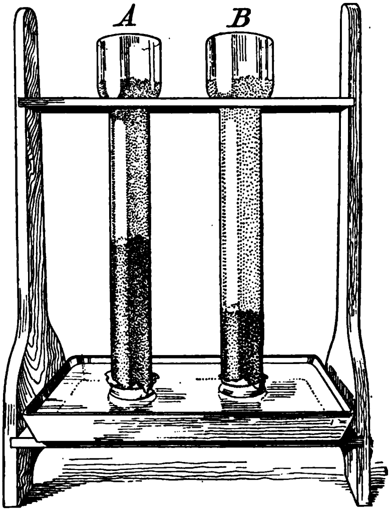

FIG. 204.— **Apparatus to illustrate the capillary rise of water in clayey and sandy soils.**

Ex. 326. Tie pieces of cheesecloth over the ends of two tall glass tubes, or two long chimneys, and fill one with a clayey soil and the other with a sandy soil (Fig. 204). Place the lower ends of the tubes in a pan of water. Does the water rise in the tubes? In which will it rise to the greater height? Explain how evaporation from the surface of the soil brings water up from below. From what water in the soil do plants get their supply? Will water in the soil move toward the plant roots? Why is film water important?

Ex. **327**. What effect does a high water table have on the root development of the plant? Will the roots of the agricultural plants penetrate below the water table? Why? Why does free water make the soil cold? What effect does a wet soil have on the plant food? How does free water affect nitrification and denitrification?

FIG. 205. — **Perforated earth-filled can showing how water leaves the soil by drainage.**

Ex. **328**. Grow two plants in tin cans, one of which has holes in the bottom while the other has not. Keep the soil in the can without the perforations saturated with water, and add the same amount of water to the other can. What difference do you note in plant growth? Explain the various ways in which underdrainage is beneficial to soils. Explain how drainage decreases the danger of injury from drought.

Ex. **329**. Make three holes in a tall can as shown in Fig. 205. Fill the can with soil and pour water on the top. Through which hole does the water escape? Is there any danger of losing good water through drainage? The apparatus shown in Fig. 206 is convenient in performing this experiment and gives conditions more nearly conforming to those in a tile-drained field.

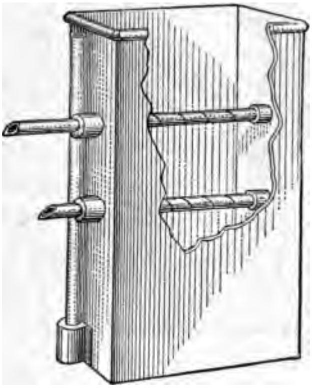

Fig. 206. — **Graham McCall drainage apparatus.**

Ex. **330.** How should you lay out a system of drainage for your home farm? Is your farm tile drained? In which kind of soil — sandy or clayey — should you lay tile deeper and farther apart? How does the water enter the tile? Why should the outlet of the tile be protected?

Ex. **331.** In what sections is irrigation most commonly practiced? Has irrigation farming any advantage over the system which depends on the natural rainfall? What are the two more common methods of applying the water?

Ex. **332.** Mix a little soda with a sandy loam soil and place it in a pot. Keep the pot standing in a pan of water. After some days note whether the soda has collected at the top of the soil. How do you explain it? Why should irrigated lands be underdrained? Remove the pot from the pan of water and pour water repeatedly on top of the soil. Does drainage remove the alkali?

Ex. **333.** What can you say about irrigation in humid climates? Do any of the farmers or gardeners in your locality use irrigation? Describe the system used by them. Are there any fields on your farm which could be irrigated at small cost?

CHAPTER LIV

TILLAGE

530. Tillage Increases Feeding Ground for Roots. Good tillage is the most efficient means of assisting nature in the conversion of unavailable plant food into forms that the plant can use. Tillage, in the sense in which it is used here, signifies any operation of stirring and pulverizing the soil by means of plows, harrows, cultivators, or any other implements, either before or after the seed is sown.

The most noticeable result of tillage is that it makes the soil finer by breaking the large lumps into smaller particles. Pulverizing the earth is beneficial in many ways. In the first place, loosening the soil makes it easier for the plant roots and root hairs to penetrate it. Mention has been made of the fact that all soils are composed of particles of rock separated by air spaces. The tender root hairs must push their way in between these soil grains, since it is impossible for them to penetrate the solid particles themselves. It must be evident that the more the soil is pulverized the larger the number of the openings between grains, and, consequently, the greater room for root growth.

Good tillage increases the amount of surface exposed to the roots by breaking the large lumps into small grains; and the more complete the pulverization, the larger the area from which the plant can obtain its food. An example will serve to illustrate what is meant. A cube 2 inches on a side presents a surface of 24 square inches. If this cube is cut once in each direction, 8 cubes are formed, each one inch on a side, giving a total of 48 square inches of surface, so that cutting only once in each direction doubles the amount of surface. Thus, theoretically, a plant should be able to derive twice as much food from the eight small cubes as from the large one.

531. Tillage Aërates the Soil. One of the most advantageous results to be obtained from tillage is the aëration of the soil. The introduction of the oxygen of the air into the soil is of benefit in many ways. It makes possible the growth of the plant roots; it enables the seeds to germinate; it aids nitrification; and it prevents denitrification.

The oxygen of the air also has a direct chemical action upon the mineral matter of the soil and tends to make it soluble. In addition it prevents the formation of certain injurious compounds, notably certain ferrous compounds.

The bacteria that enable leguminous plants to use free nitrogen are also dependent upon the air in the soil; for not only do they need oxygen, but experiments have shown that it is only from the air in the soil that they can draw their supply of nitrogen. It is necessary, therefore, in order that leguminous plants may profit by the nodule-forming bacteria, to have the soil in such condition of tilth that the air may freely circulate through it.

532. Tillage Increases Amount of Available Water. Tillage not only increases the amount of surface on which the plants can feed, but at the same time enlarges the water supply by giving the soil greater capacity for holding moisture. Attention has been called to the fact that each soil grain is surrounded by a film of water which is called capillary water or film moisture. The plant is dependent upon this film moisture for its supply, and it is readily seen that the amount of capillary water that the soil can retain depends upon the aggregate surface area presented by the particles of which it is composed.

The following quotation from King illustrates in a striking way the rate at which the film moisture in the soil increases as the soil particles decrease in size. "Suppose we take a marble exactly one inch in diameter. It will just slip inside a cube one inch on a side, and will hold a film of water 3.1416 square inches in area. But reduce the diameters of the marbles to one tenth of an inch, and at least 1000 of them will be required to fill the cubic inch, and their aggregate surface will be 31.416 square inches. If, however, the diameters of these spheres are reduced to one hundredth of an inch, then 1,000,000 of them will be required to make a cubic inch, and their total surface area will then be 314.16 square inches. Suppose again the soil particles to have a diameter of one thousandth of an inch. It will require 1,000,000,000 of them to fill completely the cubic inch, while their aggregate surface must measure 3141.6 square inches."

It will be noted that the smallest particle mentioned in the foregoing paragraph has five times the diameter of the clay particle (492). It has been estimated that a cubic foot of clay has 150,000 square feet of surface, while a cubic foot of coarse sand has 40,000 square feet of surface.

533. Tillage to Conserve Moisture. From what has been said regarding the importance of water to the plant it must be apparent that one of the chief problems of agriculture is to maintain a proper degree of moisture in the soil. It seldom happens that a crop can obtain from the soil the amount of water necessary for a maximum yield, and great skill is required to keep it from suffering for lack of moisture cube is cut once in each direction, 8 cubes are formed, each one inch on a side, giving a total of 48 square inches of surface, so that cutting only once in each direction doubles the amount of surface. Thus, theoretically, a plant should be able to derive twice as much food from the eight small cubes as from the large one.

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536. Fall plowing as well as spring plowing increases the water supply of the soil, because it leaves the ground with a loose uneven surface, in which condition it more readily absorbs the water of the winter rains. An experiment reported from Wisconsin shows that a plot plowed in the fall contained 1.15 inches more water than an adjacent plot not so plowed. Many soils in the northern states are improved by fall plowing, because the freezing and thawing brings about disintegration of the clods and leaves the soil in better tilth. The greatest advantage of fall plowing, however, is that it decreases the amount of plowing that must be done during the rush of spring work. Soils that are subject to washing cannot safely be plowed in the fall.

537. Tillage to Destroy Weeds. Weeds should not be allowed to grow, because they rob the crop of the moisture and plant food that it needs. Since weeds are usually broad-leaved, vigorous plants that transpire large quantities of water, and since there is seldom water to spare in the soil, weeds are injurious to the growing crop.

While it is probable that weeds work the greatest injury to the crop by depriving it of water, they also rob it of mineral food. Some farmers argue that if the plants remain on the ground they remove no plant food. It must be remembered, however, that they use that portion of the plant food that would be available to the crop and that the weeds must decay before this food is again rendered available. In so far as any one crop is concerned the plant food is as completely removed by a growth of weeds immediately preceding it as it would be if it were actually taken from the field. In an experiment in New Hampshire corn was grown with and without cultivation, so that in one plot the weeds were allowed to grow, while in the other the weeds were all destroyed. The first plot produced only 17 bushels to the acre, while the cultivated plot yielded 80 bushels.

FIG. 209. — A well-prepared seed bed is important in the production of large crops.

The destruction of weeds was formerly regarded as the only reason for tillage after seeding. It is now known, however, that stirring the soil has a distinct value in itself. If the farmer tills his soil, so as to reap the maximum benefits of this process, he will have no need to worry about the weeds.

538. Dry Farming. It has been estimated that one half of this country has a rainfall of less than twenty inches.

Such a rainfall is not sufficient to produce a good crop each year. In some of these semi-arid sections the land is so located that it can be irrigated, and it is being rapidly reclaimed by the irrigation projects of the national government as well as by those of private enterprise.

Fig. 210. — Barley grown three years in succession in a dry farming district. Yield 5.33 bushels per acre.

Much more of the land in these semi-arid sections, however, is beyond the reach of irrigation, and must be handled in an entirely different manner if it is to produce profitable crops. On some of these lands dry farming is followed. This is practicable on land receiving an annual rainfall of 12 to 14 inches or more, and consists in utilizing the moisture of two years to produce a crop, instead of trying to grow one each year. It is a broad application of the earth mulch.

Fig. 211. — Barley grown after fallow on land adjacent to Fig. 210. Yield 31.27 bushels per acre.

In dry farming a crop is produced in each field once in two years, so that a farmer who, for example, owns 320 acres has crops growing on 160 acres but nothing growing on the other 160

acres. Immediately after a crop is harvested the ground is plowed and harrowed to produce an earth mulch which is maintained by thorough cultivation for the remainder of that year and all through the second year. The principle of the mulch is carried still farther, for the ground after seeding, and even after the crop is up, is harrowed as long as it can be done without injury to the growing crop.

This system of tillage permits all the rain water to soak into the soil and prevents its subsequent evaporation, a condition which is very necessary because the dry air and winds of the semi-arid regions cause rapid evaporation. The soils in these regions are high in available plant food because very little of it has been leached out, and the only requisite for a good crop is an adequate supply of water.

Some of these soils settle so slowly after plowing that a special implement known as the subsurface packer has been invented for use on them. This implement packs the soil below without destroying the loose condition of the surface. In this way capillary connection is renewed between the plowed soil and that beneath, while evaporation from the surface is prevented by the mulch.

539. Summer Fallowing. A practice similar to dry farming was formerly followed in humid climates and is still advocated by some people. This practice, which is known as summer fallowing, consists in allowing the field to go through one summer without growing a crop, the surface of the soil being frequently cultivated to produce a mulch.

This method came into use at a time when the tillage implements were very crude and it was possible to control the weeds in no other way. The crop following a summer fallow is always larger than it would be otherwise, but in sections with an annual rainfall of over thirty inches, dis-

tributed throughout the year, the increase of crop seldom pays for the loss of a crop during the fallow year. There is also danger in this method of excessive loss of nitrogen by leaching. In an experiment at the New York Experiment Station over 300 pounds of nitrogen were lost in a year in the drainage of water from an acre of fallow ground, an amount equal to that removed by four 50 bushel crops of corn. No nitrogen was lost from an adjoining plot on which grass was growing, because the plants utilized the nitrates as fast as they were formed.

540. Short Fallows Desirable. Although the long summer fallow is to be recommended only when the soil has been abused and has become so foul with weeds that no other method will remove them, frequent use should be made between crops of the short fallow. It will be found advantageous in many instances to plow the land immediately after the removal of one crop and keep it well stirred until the planting of the next plot. By this means loss of moisture from the soil is prevented, the decomposition of the organic matter is hastened, and a large supply of available plant food is prepared for the succeeding crop. This method is especially to be recommended in preparation for fall-seeded crops in sections of scanty summer rainfall.

FIG. 212. — Plowing in the Orient.

541. Plows. The plow is one of the oldest tillage implements, its use antedating history. From a crooked stick (Fig. 212) drawn by an ox, a donkey, or even

a slave, the plow has developed into the steel moldboard plow (Fig. 213) of to-day. There are many kinds of plows, a kind for every purpose, the principal differences in them being the shape of the moldboard. The breaking plow (Fig. 214) has a long moldboard with a slight curve that turns the sod without breaking it. The stubble plow has a steep moldboard with a sharp, bold outward curve at its upper extremity, which gives the furrow slice a twist, causing it to break and crumble as it falls. There is a plow (Fig. 215), that has a revolving steel disk in place of the moldboard. The advantage

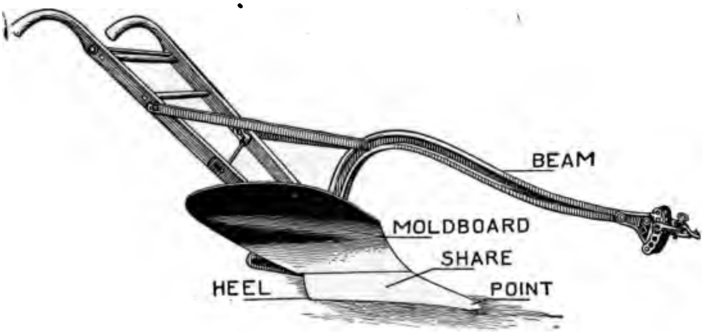

Fig. 213. — General purpose plow.

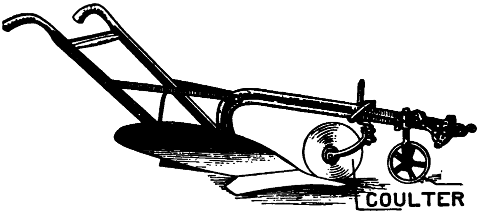

Fig. 214. — Breaking plow.

Such a rainfall is not sufficient to produce a good crop each year. In some of these semi-arid sections the land is so located that it can be irrigated, and it is being rapidly reclaimed by the irrigation projects of the national government as well as by those of private enterprise.

FIG. 210. — Barley grown three years in succession in a dry farming district. Yield 5.33 bushels per acre.

Much more of the land in these semi-arid sections, however, is beyond the reach of irrigation, and must be handled in an entirely different manner if it is to produce profitable crops. On some of these lands dry farming is followed. This is practicable on land receiving an annual rainfall of 12 to 14 inches or more, and consists in utilizing the moisture of two years to produce a crop, instead of trying to grow one each year. It is a broad application of the earth mulch.

FIG. 211. — Barley grown after fallow on land adjacent to Fig. 210. Yield 31.27 bushels per acre.

In dry farming a crop is produced in each field once in two years, so that a farmer who, for example, owns 320 acres has crops growing on 160 acres but nothing growing on the other 160

acres. Immediately after a crop is harvested the ground is plowed and harrowed to produce an earth mulch which is maintained by thorough cultivation for the remainder of that year and all through the second year. The principle of the mulch is carried still farther, for the ground after seeding, and even after the crop is up, is harrowed as long as it can be done without injury to the growing crop.

This system of tillage permits all the rain water to soak into the soil and prevents its subsequent evaporation, a condition which is very necessary because the dry air and winds of the semi-arid regions cause rapid evaporation. The soils in these regions are high in available plant food because very little of it has been leached out, and the only requisite for a good crop is an adequate supply of water.

Some of these soils settle so slowly after plowing that a special implement known as the subsurface packer has been invented for use on them. This implement packs the soil below without destroying the loose condition of the surface. In this way capillary connection is renewed between the plowed soil and that beneath, while evaporation from the surface is prevented by the mulch.

539. Summer Fallowing. A practice similar to dry farming was formerly followed in humid climates and is still advocated by some people. This practice, which is known as summer fallowing, consists in allowing the field to go through one summer without growing a crop, the surface of the soil being frequently cultivated to produce a mulch.

This method came into use at a time when the tillage implements were very crude and it was possible to control the weeds in no other way. The crop following a summer fallow is always larger than it would be otherwise, but in sections with an annual rainfall of over thirty inches, dis-

tributed throughout the year, the increase of crop seldom pays for the loss of a crop during the fallow year. There is also danger in this method of excessive loss of nitrogen by leaching. In an experiment at the New York Experiment Station over 300 pounds of nitrogen were lost in a year in the drainage of water from an acre of fallow ground, an amount equal to that removed by four 50 bushel crops of corn. No nitrogen was lost from an adjoining plot on which grass was growing, because the plants utilized the nitrates as fast as they were formed.

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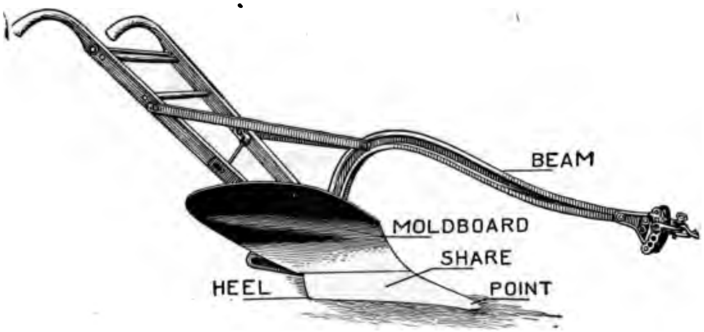

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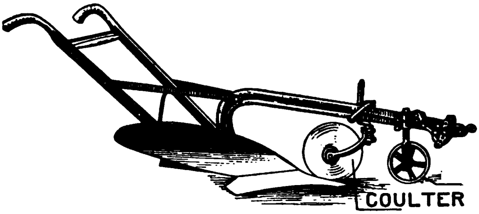

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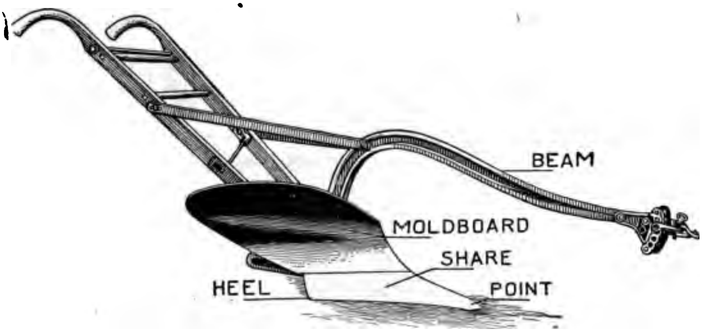

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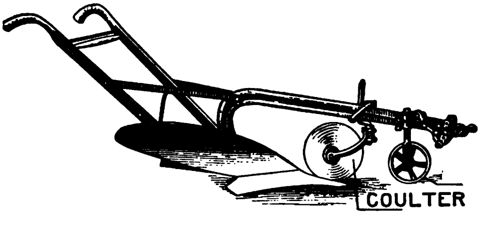

Fig. 214. — Breaking plow.

the perspiration of the hands is usually acid in reaction. The safest way is to use a pair of small forceps in handling the paper. Only a sensitive litmus paper, such as is used in the chemical laboratory, should be employed in making the test. The cheap blue litmus papers so often found in the drug stores are strongly alkaline and will not change color even in soil which is sour enough to prevent entirely the growth of clover.

552. The Truog test for acidity consists in mixing the soil to be tested with a small quantity of calcium chloride and a very little zinc sulphide. Water is added and the mixture is heated to boiling. A strip of lead acetate paper is held over the mouth of the flask for two minutes while the boiling proceeds. If the soil is acid it reacts on the zinc sulphide and forms hydrogen sulphide, which darkens the lead acetate paper (87). If no acid is present in the soil no darkening of the paper will occur. When the directions are carefully followed the color of the lead paper gives some idea of the degree of acidity in the soil. A small outfit including an alcohol lamp for heating water so that the test may be used in the field is now on the market (Fig. 226).

FIG. 226. — **Apparatus for Truog test.**

553. Hydrochloric acid is used to show the presence of an abundance of limestone in the soil. A handful of soil is pressed into a ball, and then a depression is made in it to hold a little dilute hydrochloric acid. If a considerable amount of limestone is present, it will be shown by effervescence caused by the action of the acid on the calcium carbonate. A decided effervescence indicates that the soil contains sufficient limestone for all purposes. The small amount of air which is forced out of the soil by the liquid must not be mistaken for effervescence from the carbonates. This test is of limited application only, since a soil may be neutral in reaction and able to produce clover and yet not contain enough carbonate to give a noticeable evolution of carbon dioxide.

554. Correcting Acidity. When any of the tests indicate that the soil is sour, it is necessary to add to it some substance that will neutralize and thus destroy the acid. The cheapest material that can be used for this purpose is ordinary limestone. The stone is pulverized by crushers (Fig. 227) and rollers until it is fine enough to be applied with a lime drill. A powder that will all pass through a sieve having ten meshes to the linear inch is sufficiently fine, although the finer the stone the quicker it will act on the soil.

Fig. 227. — A small portable pulverizer for grinding limestone.

The ordinary application of ground limestone to acid soils is two tons to the acre. It is best applied by means of especially devised lime spreaders (Fig. 228) which can be set to distribute from five hundred pounds to four or five tons to the acre. A manure spreader may be used by covering the bottom with straw and placing the powdered limestone on top, or a skilled worker may even spread it from the wagon by means of a shovel. If much limestone is to be used, the special spreader will soon pay for itself.

FIG. 228. — Applying ground limestone with a spreader.

It sometimes happens that a farm that has an acid soil also has outcroppings of limestone. Small portable limestone crushers are on the market at a low price, which make it possible under such conditions for the farmer to crush his own limestone at a small expense. It is to be noted that even in limestone sections acid areas are frequently found, especially on the high lands that have been under cultivation for a long period.

555. Other Forms of Lime. Although under ordinary circumstances ground limestone is the cheapest and best material to use for correcting the acidity of soils, it may happen that the land is so far removed from any source of limestone that the freight charges are excessive. In such a case it might be more economical to use quicklime (CaO) in place of the limestone. In burning 100 pounds of calcium carbonate 56 pounds of lime are produced; but the latter has the same power to neutralize acids as the former. Thus it will be seen that not much more than one ton of quicklime will do the same work as two tons of limestone.

Hydrated lime, which is merely another name for slaked lime ($Ca(OH)_2$), is often recommended for use on the soil. It is effective in neutralizing acidity, but is usually too high in price for that purpose.

Air-slaked lime (115) is formed when quicklime is exposed to the air. If it stands for a long time the calcium is all converted into the carbonate, with the result that the material is practically the same as very finely ground limestone. One ton of quicklime is equivalent in neutralizing power to 2640 pounds of hydrated lime, or 3570 pounds of carbonate of lime; and these relations should be kept in mind when determining which substance is most economical to buy.

556. Marl. In many places beds of marl of considerable size are found. Most of the marls are formed from shell deposits, and consist of carbonate of lime of more or less purity. As marl is practically the same as ground limestone, it has the same effect upon the soil, and is a convenient form in which to use lime when obtainable at reasonable cost. Some of the European marls contain appreciable quantities of potash and phosphoric acid as well, but the American marls are of value only for the lime they contain.

557. When to Apply Limestone. The principal reason for using limestone or the other forms of lime is to enable the soil to produce clover or other legumes; consequently the limestone should be used far enough in advance of the leguminous crop to insure the neutralizing of the acidity of the soil before the legume is planted. If corn is grown in the rotation, it is advisable to spread the limestone on the ground immediately after plowing, and to harrow it in. In this way the limestone is thoroughly mixed with the soil and has time to destroy the acids before the legume is seeded.

558. Field Test with Limestone. The tests for acidity already described are all good in their way; but the importance of having right conditions for the growth of clover is so great that a practical test of limestone in the field should be made in every case where the growth of clover is unsatisfactory. A strip across the field should be dressed with limestone at the rate of forty pounds to the square rod, and its effect upon the growth of clover noted.

559. Lime Improves the Physical Condition of Soil. Lime also has a very marked effect on the physical condition of the soil. When added to a sandy soil it tends to make the soil more compact by partially cementing together the particles of sand and thus making the soil capable of retaining larger quantities of water. When used on clay lands, on the other hand, lime makes the soil more mellow. A clay soil containing very little lime is made fine with the greatest difficulty; it adheres to the implements used when wet, and cracks when allowed to dry. A soil rich in lime crumbles more easily than one lacking it, is readily brought into good tilth, and does not adhere to any appreciable extent to the implements. The addition of lime to a soil containing much clay makes the soil more friable, makes it possible for the rains to percolate more easily through the soil, and overcomes the danger of puddling. The puddling of clay soils is due to the fact that the clay is composed of very small granules which fit so closely together that the water cannot pass between. When lime is added to the soil a number of these small particles become cemented together to form a much larger granule, and as the granules increase in size the spaces between them also become larger.

This effect of lime may readily be shown by taking a sample of clay, adding a little water, working it thoroughly, and then allowing it to dry. As a result of this treatment it becomes as hard as a brick. If to another portion of the clay a little lime is added (say one per cent) and this is moistened, mixed thoroughly, and allowed to dry, it will be found that a mere touch will cause it to crumble to pieces. This granulated condition of clay soils, so easily accomplished by liming, is not readily destroyed but will last for some years. This mechanical effect on soils is more marked in the case of quicklime than it is with limestone.

560. Lime Makes Potential Food Available. Lime is useful in liberating the unavailable food materials of the soil. Most of the potassium of the soil, for instance, is locked up in insoluble silicates and is not available to plants. Lime decomposes some of these silicates and converts the potassium into forms that the crop can use. Experiments have proved that when lime is applied to a soil originally poor in this substance, the plants grown are not only richer in lime but also in potash. The use of lime then may for a time have a similar effect to that of potash-containing manures, but it must be remembered that the lime does not supply potash; it merely makes available the potash that is present in the soil; and if the store of potash originally present is small, it will probably need liberal potash manuring at an earlier date because of liming.

Lime is also beneficial in preventing the formation of the very insoluble iron and aluminum phosphates, or in changing them to calcium phosphate if they are already formed. Another effect of adding lime or limestone to an acid soil is to make the phosphates of the soil available.

561. Lime not a Universal Remedy. So much has been written about the use of lime that there is danger of creating the impression that lime or limestone is the universal remedy for all unproductive soils, and that no other treatment than liming is necessary. It must be remembered that lime adds no plant food save calcium to the soil, but simply brings about conditions that enable the crop to use larger quantities of the food already present, so that if used alone it makes the exhaustion of the soil more rapid. Lime can in no way take the place of good tillage, drainage, manure, or fertilizers. There is an old saying that " lime makes the father rich, but the son poor," and this is undoubtedly true if lime is used alone. It has, however, a legitimate place in agriculture, and if used in connection with green crops, barnyard manure, and commercial fertilizers will in many cases produce beneficial results.

562. Acid Resistant Plants. Not all crops are seriously injured by an acid condition of the soil. Such plants as the cranberry and blueberry actually require an acid soil and will not grow on one that is alkaline. Other crops, while not preferring an acid soil, are injured but little by acidity. The potato may be grown very successfully on an acid soil and under such conditions is less subject to the attacks of the scab fungus than when grown in neutral or alkaline soils. The potato scab fungus grows more readily in an alkaline soil; consequently lime should never be added to the soil immediately preceding the potato crop.

EXERCISES

Ex. 343. What is meant by a sour soil? What is the cause of acidity in soils? What effect does soil acidity have on plant growth? What plants are especially sensitive to acidity? What effect does acidity have on the fixation of nitrogen? How does acidity affect nitrification? Why is limestone necessary for nitrification?

Ex. 344. Explain how the character of the vegetation indicates whether or not soil is acid. See if you can find sorrel growing abundantly anywhere in the vicinity of the school. What should you suspect regarding the soil when clover fails?

Ex. 345. Test a number of soils with litmus paper (551). Are any of the soils acid? Mix a little limestone thoroughly with the soil and make the litmus paper test again on the following day. Give the result. Why should the litmus paper never be handled with perspiring fingers? Are there any acid soils on your home farm?

Ex. 346. Test a soil for acidity by the Truog test as follows: (1) Prepare the lead acetate paper by dipping a piece of white filter paper into a 10 per cent solution of lead acetate. Spread the paper on a pane of glass to dry. (2) Prepare the zinc sulphide and calcium chloride mixture by dissolving 50 grams of neutral calcium chloride in 250 cc. of water, and then adding 5 grams of finely pulverized zinc sulphide. (3) To perform the test place 10 grams of soil in a boiling flask, adding 5 cc. of the zinc sulphide-calcium chloride mixture and 95 cc. of distilled water. Heat to boiling and boil one minute, then place a strip of the lead acetate paper over the mouth of the flask, and boil two minutes. A darkening of the acetate paper indicates the presence of acid, the degree of acidity being approximately shown by the depth of the color. What causes the paper to darken? Why does it not change color if the soil is neutral? Why is it essential that the calcium chloride should be neutral?

Ex. 347. Test some soils for limestone (553). Is limestone present? Is this test of general application? Would there be any need of liming a soil that showed the presence of limestone by this test?

Ex. 348. What method is used to correct the acidity of soils? How much limestone is ordinarily used to the acre? How finely should the limestone be pulverized? How may it be applied to the soil?

What other forms of lime are used to correct acidity? What is marl? Is it suitable for use on the soil?

Ex. **349.** What is the principal reason for using limestone? Why should it be used in advance of clover? Why is it advisable to apply it to the ground that has been plowed for corn? How should you conduct a field test with limestone? Is such a test advisable?

Ex. **350.** Take a handful of wet clay and thoroughly knead it into a ball and dry it in the oven. Add one per cent of lime to another handful of the clay and knead it into a ball and dry it in the oven. How do the two samples behave when broken with a hammer? What effect does lime have on the physical condition of soils?

Ex. **351.** Place a tablespoonful of clay soil in each of two tall glass cylinders or bottles. Shake the cylinders to get the clay in suspension. To one cylinder add a little slaked lime and stir. What difference is there in the behavior of the clay in the two cylinders? How do you account for it? Why do clays puddle? How does lime prevent puddling?

Ex. **352.** What effect does lime have on the potential plant food of the soil? Why do acid soils seem to contain more phosphorus after liming? In what form is the phosphate in acid soils? If a farmer uses lime alone, what effect will it eventually have on the fertility of his land?

Ex. **353.** Are all plants injured by acidity? Are there any that prefer an acid soil? Why should care be used in liming potato soils?

CHAPTER LVI

ORGANIC MATTER

563. Old and New Soils Compared. If a soil that has grown crops continuously for many years without the addition of plant food is compared with an adjacent plot of virgin soil, a marked difference will be found in the amounts of organic matter that the two soils contain. Soils that have been under cultivation for periods of twenty to thirty years with no provision for maintaining the organic matter are found to contain less than two thirds as much organic matter as the original soil. There can be no doubt that, in a great many instances, the loss of fertility is due to the rapid decrease in the amount of organic matter.

564. Organic matter in soils is largely derived from the remains of the plants that have grown on them. Under natural conditions the entire plant becomes a part of the soil and furnishes it with an abundance of organic matter. When the crop is removed from the field, only the stubble and roots remain to supply the vegetable matter, but the quantity supplied in this way is not sufficient for the best results. Some of the organic matter of the soil is of animal origin, but the amount so derived is insignificant as compared with that derived from vegetation.

Organic matter exists in the soil in all stages of decomposition. Decay begins as soon as the plants die or other organic matter is added to the soil, and is due to the action of the bacteria that are always present in the soil. The soil contains, therefore, fresh vegetable material that has not begun to decay; partially decayed material that still retains a part at least of the original form; and organic matter that has so completely decomposed that it has entirely lost the physical structure of the material from which it was derived. The black waxy material coating the soil grains, giving the dark color to some soils, is organic matter in a very advanced state of decomposition, and is known as *humus.* The term humus is often used as a synonym of the term *organic matter.* Such usage, however, is incorrect, as much more than the humus is included in organic matter. All forms of organic matter are of importance in the soil; but probably that part which is undergoing active decay is the most beneficial.

565. Organic Matter Increases the Amount of Soil Water. Organic matter has a high absorptive power for water. A sponge, for instance, which is a good example of organic matter, will absorb and retain more than ten times its own weight of water. Cellulose and other forms of vegetable matter will hold practically the same amount of water. It will readily be seen, therefore, that the more organic matter the soil contains the greater will be its power to store water. In one experiment the addition of one per cent of organic matter to the soil increased its water-holding capacity a little over eight per cent. The water retained by the organic matter is in large part so loosely held that it can be utilized by growing plants. The following table gives the amount of water held in a cubic foot of three different soils with varying amounts of organic matter:

KIND OF SOIL	POUNDS OF WATER IN ONE CUBIC FOOT
Sand	27.3
Sandy loam	38.8
Loam	41.4

It will be seen that the quantity of water increases with the amount of organic matter present, the sand containing the least, and the loam, which has the largest percentage of organic matter, containing much more. It is evident from the above that one of the best ways to enable the soil to store water from the spring rains for the use of the growing crop during the hot summer months is to keep the soil well supplied with organic matter.

566. Organic Matter a Storehouse for Plant Food. The organic matter of the soil contains part of the plant food that was utilized by plants formerly grown on the soil. It has been shown that most of the nitrogen in the soil is present in the organic matter and that this becomes available through the gradual decay of the plant. It is fortunate that most of the nitrogen is stored in organic matter; for if it were all immediately converted into nitrates, there would be great loss by leaching, since the nitrates are very soluble and the soil has little power of retaining them.

567. Organic Matter Makes Potential Food Available. The presence of decomposing organic matter in the soil is an important factor in making the mineral elements of plant food available. During decay certain acids, such as lactic, acetic, and nitrous acids, are produced, and these undoubtedly have a solvent action on the mineral matters of the soil, tending to make them more available to the plant. Perhaps quite as important a factor is the large amount of carbonic acid formed during the process of decay. This carbonic acid dissolved in the soil water is of prime importance in the production of soluble plant food, and it also has a beneficial effect on the physical condition of the soil, especially if the soil contains a large amount of clay.

568. Organic Matter Improves Soil Texture. Organic matter is also valuable in improving the physical condition of the soil. Sandy soils are made more compact by its presence and better able to supply a crop with food and moisture. Clay soils, on the other hand, are made more mellow by the addition of organic matter. Clay is likely to become too compact unless there is a certain amount of organic matter present to prevent it. The better tilth of a soil due to the presence of organic matter facilitates drainage and ventilation, both of which are necessary to the promotion of soil sanitation.

Organic matter is an essential part of all true soils. In most soils there is a marked difference in color and texture between the surface, or true, soil and the subsoil. The surface soil is darker in color, less compact, and more easily worked by tillage implements than is the subsoil. The difference is due largely to the greater amount of organic matter contained in the surface soil. The darker color of the soil is due largely to humus, the very black soils containing large quantities of humus. Organic matter, more especially humus, also affects the temperature of the soil, for the darker-colored soils absorb more heat from the sun's rays than do the lighter-colored ones.

569. Loss of Organic Matter. Investigations have shown conclusively that as the organic content of the soil is decreased by constant cultivation and cropping, the nitrogen content of the soil, the amount of moisture that it contains, and the crop production are likewise decreased.

All the methods so far discussed for making potential plant food available tend to decrease the amount of organic matter in the soil. Tillage, drainage, bare fallowing, and liming the soil all increase the amount of food available to the crop, because they present ideal conditions for the decomposition of organic matter in the soil; but dependence upon these methods alone will eventually result in injury through loss of organic matter. The loss of nitrogen and organic matter is strikingly shown in the following table:

	Native Soil Per Cent	Cultivated 23 Years Per Cent	Loss Per Cent
Organic matter	3.97	2.59	1.38
Nitrogen	0.36	0.19	0.17
Capacity to hold water .	62.00	54.00	8.00

The foregoing statements should not be construed as arguments against tillage, drainage, and liming, because the destruction of organic matter is an essential part of good farming. It is the farmer's business to bring about in the soil the proper conditions for the rapid decay of organic matter so that the crops may utilize the plant food therein contained. However, if he is to have continued success in producing crops, the farmer must at the same time return organic matter to replace that which has been destroyed.

570. Restoring Organic Matter. Under farm conditions where most of the crops must be removed from the field, the maintenance of the supply of organic matter is a serious problem. A certain amount of organic matter is left behind in the stubble and the roots. This material should be utilized to the fullest extent. The practice of burning over the field to destroy the stubble before plowing is to be condemned, because large quantities of organic matter are destroyed in this way. The straw, cornstalks, and any similar material, if not used for feeding, should be incorporated into the soil. The burning of any of the organic matter of the farm or garden is never justified except when necessary to prevent the spread of plant diseases.

Large amounts of organic matter may be restored to the soil by a careful use of the stable manure. The manure from the domestic animals contains nearly one half of the organic matter of the feeds, the rest having been oxidized in the animal body and given off largely as carbon dioxide and water. If the manure is properly handled, therefore, about one half of the organic matter in all feeds used may be added to the soil. The precautions necessary to obtain these results are discussed in Chapter LVIII.

The plowing under of grass sods adds organic matter to the soil. Such grasses as Kentucky blue grass, especially, fill the soil full of small, fibrous roots and when the sod is plowed the roots and the stubble begin to decay. In one instance the first six inches of soil in a field of timothy and redtop were found to contain nearly four tons of roots and stubble to the acre. Pasture lands, if well fertilized so as to grow an abundance of grass, may be used as a means of maintaining the organic matter of the farm.

571. Green Manuring. On many farms even the complete utilization of the various methods outlined in the last section will not suffice to maintain the necessary amount of organic matter in the soil for the production of maximum crops. Under such conditions it is necessary to grow a crop for the express purpose of plowing it under to increase the organic matter of the soil. Such a practice is known as *green manuring*. Plowing under green crops raised for that purpose is one of the oldest means of improving the fertility of the soil. It was advocated by Roman writers more than two thousand years ago, and has been in more or less common use among progressive farmers ever since.

FIG. 229. — Plowing under a green manure crop.

The value of green manuring depends primarily on the fact that it increases the amount of organic matter in the soil. Several kinds of crops may be used as green manures, but the most valuable for this purpose are the legumes. The discovery that the leguminous plants can, through the nodule-forming bacteria, fix the free nitrogen of the air, has thrown a new light on green manuring and the plants adapted to this purpose. The legumes have all the advantages of the other plants in providing organic matter, and at the same time they increase the amount of nitrogen in the soil. They are, as a rule, deeper-rooted plants and are supposed to bring up mineral food from the subsoil, and leave it where it will be within reach of the more shallow-rooted plants. Of the legumes the crops most often recommended are red clover, sweet clover, cowpea, crimson clover, the lupines, soy bean, and the ordinary field bean, and field pea. Of these, red clover is probably the one most generally used.

572. Green Manuring and Type of Farming. Such crops as red clover, which make the best green manures, also have great value as feeds for live stock, and it may be found more profitable to feed them to animals and return the manure to the soil than it is to turn them under. But even on stock farms it is often advisable to plow under the second growth of clover instead of cutting it for hay.

In any system of farming in which the crops are sold from the farm, some provision for green crops to plow under is absolutely necessary, and the rotation used should include a green manure crop. The increase in the other crops from this practice will more than make up for the fact that there is no crop to sell from the green manure field.

573. Catch Crops for Green Manuring. Where it is inadvisable to devote an entire season to the growth of a crop for green manuring, good results may often be obtained by the use of what is known as a catch crop, or a crop grown between two main crops. Rye is often planted in the corn land at the time of the last cultivation and allowed to grow until the ground is plowed the following spring, thus adding organic matter to the soil. In the southern states crimson clover and other legumes are used in a like manner, but in the north the legumes are uncertain as catch crops. A mixture of rye and hairy vetch is very satisfactory as a catch crop after corn. The use of cover crops in orchards is another example of a catch crop.

FIG. 230. — Crimson clover as an orchard cover crop.

574. Danger from Green Manuring. While green manuring is a valuable method of increasing the humus supply of the soil, it is not unattended by danger. In a dry season, for instance, the growth of a crop to plow under may result in lowering the moisture content of the soil to a point that is detrimental to the succeeding crop. There is also danger in such a season that there may not be sufficient moisture in the soil to bring about the decomposition of the organic matter that is turned under, the result being serious injury to the physical condition of the soil. Such injury, however, does not frequently occur, and its bad effects are only temporary. If a crop is plowed under during a dry season, the ground should be rolled with a heavy roller so as to renew the capillary movement of moisture between the surface and the subsoil.

EXERCISES

Ex. **354.** If possible, obtain a sample of soil from the center of a field that has been under cultivation for a long time, and another sample from the fence row of the same field. What difference do you note in the two samples? Does the soil from the fence row contain more organic matter than the other? What is the source of the organic matter in soils? What is meant by humus? Is humus identical with organic matter? What gives the black color to some soils?

Ex. **355.** Weigh a large sponge after it has been thoroughly dried. Now dip the sponge in water, hold it up until dripping has ceased, and weigh it again. How many times its own weight of water will the sponge hold? What can you say about the power of organic matter to increase the water-holding capacity of the soil?

Ex. **356.** Determine the water-holding capacity of a soil as described in Ex. 318. Add to another sample of the soil one per cent of ground moss or wheat bran and determine the water-holding capacity. How much does the organic matter increase the water in the soil?

Ex. **357.** Explain what is meant by the statement that organic matter is a storehouse of plant food. How is this plant food made available? How does organic matter make the mineral food of the soil available?

Ex. **358**. Note the difference in appearance between the surface soil and the subsoil. What makes the surface soil darker in color and more friable than the subsoil? What effect does organic matter have on the texture of a sandy soil? Of a clay soil?

Ex. **359**. Explain how organic matter is lost from the soil. Show how some of it may be restored by the plant residues, by the use of stable manure, and by plowing under sods.

Ex. **360**. What is meant by green manuring? Is it a modern practice? What advantages have the legumes as green manure crops? What are the crops most commonly used as green manures? In what type of farming is green manuring desirable?

Ex. **361**. What are catch crops and how may they be used to increase the organic matter of the soil? Give an example of the use of catch crops. Are there any dangers connected with green manuring? Why should the roller be used when a heavy crop is plowed under?

CHAPTER LVII

ROTATION OF CROPS

575. Origin of Rotations. It is the common knowledge of farmers in those parts of the world where the land has been cultivated for a long time that the fertility of the soil is maintained for a much longer time by growing a variety of crops than by producing one crop continuously. The adoption of a system of rotation of crops has been the outgrowth of accident rather than the result of an understanding of its underlying principles. The system of alternating years of bare-fallow and wheat may be said to be a two-year rotation and was the first to be adopted. History teaches us that this was later followed by a three-year rotation consisting of fallow, wheat, beans, or oats; and still later, when the value of clover and fallow crops became evident, this rotation gave way to the now famous Norfolk rotation of turnips, barley, clover, and wheat, the typical English rotation. The Norfolk four-year course represents the more common type the world over, consisting as it does of cereals alternating with hoed crops and leguminous crops.

576. Plants Differ in Food Requirements. There are many arguments to be advanced in favor of growing a variety of crops on the soil. The different crops vary in their food requirements and in their ability to procure this food from the soil. When one crop is grown continuously on the same field, nearly all the plant food that it finds available may become exhausted, although the soil will still contain large quantities of food in forms that could be assimilated by plants of another class. Some crops evidently require the mineral matter to be in a readily soluble form, while others can use less available forms of plant food. Other crops make an especial drain on one element of plant food. By growing plants with different food requirements the different elements are more evenly used, and there is less likelihood of any one element becoming exhausted.

577. Plants Differ in Manner of Growth. The various crops differ widely in their systems of root growth. Some plants, as wheat, for example, are comparatively shallow-rooted and must obtain their food from the surface soil. Others, as the clovers, are very deep-rooted and are able to use food that is not within the reach of the more shallow-rooted plants. The deep-rooted plants are not only able to procure the low-lying food, but probably bring a part of it to the surface, where it remains, upon their decay, for the use of the succeeding crop. It is well known that the shallow-rooted plants do better when preceded by a deep-rooted crop.

578. Rotation Improves the Soil and Economizes Labor. When plants of different varieties are grown, the soil receives different treatment for each crop; so that the faults of one year are likely to be corrected the next year. Thereby the soil is kept in much better physical condition. As a general rule the ground can be better prepared for the succeeding crop if a judicious rotation is practiced than if the same crop is grown continuously. The roots and stubble of clover and grasses are also factors of some importance in improving the texture of the soil. Everything considered, the tilth of the soil will be found to be much improved by rotation.

The growing of a variety of crops on the farm results in economy of labor; for the work of caring for them is distributed throughout the season instead of all coming at one time. In this way it is possible to secure cheaper and better help than when only a few kinds of plants are raised.

579. Rotation Aids in Controlling Diseases, Insects, and Weeds. Rotation also enables the farmer to control plant diseases and to head off the injurious insects. Most of the plant diseases are caused by bacteria or fungi that live only on one genus of plants, or, at any rate, are more or less restricted in the number of crops that they can use as host plants. Where one crop is grown continuously, these disease-producing fungi have every opportunity to be carried over from one year to another. Most of the injurious organisms are comparatively short-lived, so that if three or four years of crops that are not suitable host plants intervene, these organisms are likely to be destroyed.

In the same way it may be said that the injurious insects are limited to certain plants for their food supply, and if these plants are not grown on the field for a number of years, the insects may die from starvation. These remarks do not apply, of course, to those insects that have migratory powers. There is no doubt that both diseases and insects can be more easily suppressed if rotation is practiced.

Where one crop is grown continuously, the soil becomes infested with certain weeds that are not destroyed by the system of tillage necessary for that crop. The varying treatment to which a soil is subjected in a well-planned rotation makes this condition impossible; so that the destruction of weeds may be considered as one of the very desirable results of a rotation of crops. In lands badly infested with particular weeds it may even be desirable to omit from the rotation for a while the crop whose growth presents the best condition for the propagation of these weeds.

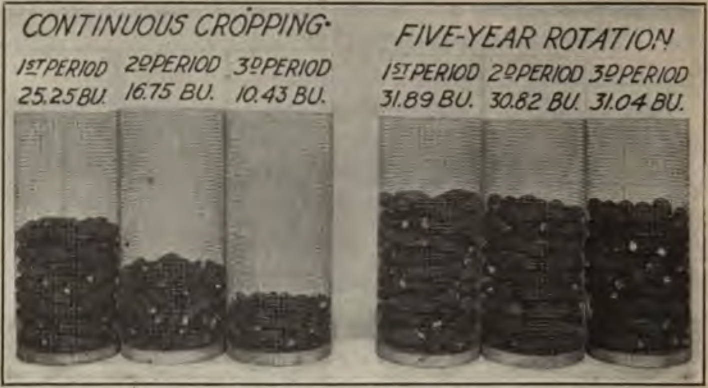

Fig. 231. — Showing the advantage of rotation in corn production.

580. Rotation Increases Crop Yields. When crops are grown in rotation, each crop gives a higher yield than when it is grown continuously on the same field. This is true whether the crops are grown with or without manures or fertilizers. The effect of rotation upon the yield is shown in the following table, which gives the average yield for the third five-year period of three different crops grown continuously, and in a five-year rotation at the Ohio Experiment Station.

Crop	Unfertilized	Stable Manure	Commercial Fertilizers
Corn			
Grown continuously . .	10 bu.	24 bu.	39 bu.
Grown in rotation . . .	31 bu.	50 bu.	54 bu.
Oats			
Grown continuously . .	22 bu.	35 bu.	45 bu.
Grown in rotation . . .	33 bu.	42 bu.	53 bu.
Wheat			
Grown continuously . .	6 bu.	12 bu.	17 bu.
Grown in rotation . . .	14 bu.	25 bu.	33 bu.

581. Planning a Rotation. In planning a rotation the farmer must be guided by his own conditions and his requirements in the way of crops. A rotation for a dairy farm, for example, might be quite different from one for a farm devoted to the production of grain for the market. A few general rules, however, will apply to all rotations. Every rotation should include one intertilled or hoed crop, such as corn, potatoes, or cotton, in order that the soil may receive the benefit of such a crop in the way of destroying weeds, improving the tilth, and setting free potential plant food. At least one leguminous crop should be included. A crop that is exacting in its food requirements should be followed by one less exacting. In general terms, the crops should vary as much as possible in their food requirements, manner of growth, root systems, and the season of the year in which they occupy the ground. Whenever possible, the rotation should include a catch crop or provide some other method of insuring an adequate supply of organic matter. Fertilizers should be applied to the particular crop or crops that will give the most profitable returns for their use.

582. Some Typical Rotations. Rotations are in use that cover periods varying from two to seventeen years. In general, short rotations of three to five years are more in favor than very long ones. A few examples will exemplify the principles of rotation.

A common five-year rotation is as follows: Corn, oats, wheat, clover, and timothy, each one year. Rye may be seeded as a catch crop in the corn.

Corn, wheat, and clover form a common three-year rotation, the ground being plowed for the corn, and the wheat seeded in the corn stubble after disking.

Potatoes, wheat, and clover are popular as a rotation in some of the good potato sections. The clover is plowed under for potatoes and a heavy application of fertilizers is used.

A rotation often used on dairy farms consists of corn, oats, and clover, besides timothy for pasture. A catch crop may be used in the corn.

A rotation that has been recommended for grain farming in the Middle West is corn, oats, clover, and wheat. The clover is plowed for wheat and all the straw and cornstalks are plowed under. A catch crop of clover is also grown between the wheat and corn. Ground rock phosphate (609) is plowed under with the clover.

A rotation advocated for the cotton-growing sections consists of cotton, corn, oats, and cowpeas, the last named crop being plowed under as a green manure.

Over forty different rotations are in use in one state, many of which, however, do not meet the requirements suggested above for planning the rotation.

EXERCISES

Ex. **362.** Explain the evolution of the rotation of crops. Is rotation of crops desirable? Outline the reasons in favor of rotation under these headings: (1) Food requirements of plants. (2) Manner of growth. (3) Effect upon soil and economy of labor. (4) Effect upon diseases, insects, and weeds.

Ex. **363.** Are the yields of the various crops increased by rotating? Give the general rules for planning a rotation. What rotation is used on your home farm? Do you use catch crops? Do you grow any crop to plow under? How many different rotations can you find in use in your community? How do they accord with the suggestions for planning a rotation?

CHAPTER LVIII

STABLE MANURE

583. If the crop yields are not to decrease, some way must be provided to replace in the soil the plant food removed by the crops. This is done by the use of stable manure and commercial fertilizers.

584. Importance of Stable Manure. Stable manure is the oldest and is still the most used of all fertilizers. It has stood the test of long experience, and has proved its position as one of the most important manures. The fact that the application of the excrement of animals to the soil results in increased crop production is mentioned by the early Roman writers, and from that time to the present the majority of farmers have placed their main reliance on this class of manures for maintaining fertility of the land.

The importance of manure is shown by the fact that the quantity produced annually by the domestic animals of the United States contains an amount of nitrogen, phosphorus, and potassium that would cost $2,458,470,000 if purchased in the cheapest forms of commercial fertilizers. More than one third of the value of the manure is lost through improper handling, and this loss is replaced only in small part by the one hundred million dollars' worth of commercial fertilizers purchased annually by the farmers of this country. This loss of plant food is the more unfortunate because it could in great measure be prevented.

585. Valuation of Manure and Fertilizers. Since some way of stating the value of manure is desirable, it is customary to use the same method that is employed in calculating the value of commercial fertilizers; namely, to base the valuation on the quantities of nitrogen, phosphorus, and potassium that the manure or fertilizer contains, ignoring any other constituents that may be present. The fertilizer trade has always stated phosphorus not as the element but as the anhydride, P_2O_5, which is termed *phosphoric acid* in the trade (179). Likewise potassium is stated as the oxide, K_2O, under the trade name of potash. One pound of phosphorus is equivalent to 2.3 pounds of phosphoric acid (P_2O_5). One pound of potassium is equivalent to 1.2 pounds of potash (K_2O). Nitrogen is sometimes stated as ammonia (NH_3), but this term is not so commonly used as are the terms phosphoric acid and potash. One pound of nitrogen is the equivalent of 1.2 pounds of ammonia. As these names are used almost exclusively in the fertilizer trade, and most commonly in the writings on manures and fertilizers, and are also recognized by law in nearly all the states, it seems best to employ them in this text. The term phosphoric acid as used in discussing fertilizers does not mean true phosphoric acid (H_3PO_4) (175) nor does the potash of the fertilizer trade mean true potash (K_2CO_3) (204).

For purposes of valuation nitrogen is given a price of 15 cents a pound, and phosphoric acid and potash are quoted at 5 cents a pound each. These are average prices for the past decade and will be used in all the calculations in this text.

586. Composition of Manure from Different Animals. The manures produced by the various classes of animals differ in their composition and in their physical properties. The following table gives the amount and value of the plant food in one ton of manure of the common domestic animals.

ANIMAL	NITROGEN	PHOSPHORIC ACID P_2O_5	POTASH K_2O	VALUE PER TON
Cow	9 lb.	3 lb.	8 lb.	$1.89
Pig	9 lb.	4 lb.	12 lb.	2.14
Horse	12 lb.	6 lb.	11 lb.	2.55
Sheep	17 lb.	5 lb.	13 lb.	3.39
Chicken	28 lb.	18 lb.	7 lb.	5.44

The difference in the value per ton of the manures is largely due to the varying amounts of water which they contain. Chicken manure, sheep manure, and horse manure contain less water than the manure from pigs and cows.

While the table shows that there is a decided difference in the value of a ton of the different manures, it is also true that the value of the total amount of manure produced from the same feeding stuffs does not vary much for the different animals. If the same kinds and amounts of feeds were given to cows and to sheep, for instance, the cows would produce more tons of a wetter manure than the sheep; but the total value would be about the same in either case.

587. Factors Affecting the Value of Fresh Manure. The value of stable manure as produced, and before it has been subjected to any of the losses to be discussed later, is largely determined by five principal factors: namely, (1) the kind of feeds, (2) the age of the animal, (3) the kind of animal, (4) the products from the animal, (5) the kind and amount of litter used.

588. Kind of Feeds. The total value of the manure produced by a given live weight of animals depends upon the quality and the quantity of the feeding stuffs used in the ration. Feeds vary greatly in the amount of plant food that they contain. The fertilizing value of a ton of timothy hay, for example, is $5.21; of clover hay, $8.79; of wheat bran, $12.52; and of cottonseed meal, $23.20. Animals fed on food substances low in fertilizing value will produce manure of a correspondingly low value. In one experiment with two lots of pigs, one of which was fed on corn meal and bran, and the other on corn meal and meat scraps, the manure produced by the latter had twice the value of that from the pigs fed corn and bran.

589. Age of the Animal. Young animals use some of the nitrogen compounds of the ration to build their muscles, and a part of the phosphoric acid and calcium is utilized in forming bone. Mature animals, whose bones and muscles are already developed, retain in their bodies very little of the fertilizing constituents of the feeds. Manure from young and growing animals, therefore, has a lower value than that from mature animals.

590. Kind of Animal. The kind of animal, as has been stated, affects the value per ton of the manure more than it does the total value of the manure produced from the same feeds. Pigs and cows, however, consume more food in proportion to their weight than do horses and sheep, and, consequently, produce manure of a greater total value during the year.

591. Product of the Animal. Milch cows use a portion of the nitrogen and phosphorus of the feeds in producing milk, and in that way some of the fertility value of the manure is lost. The value of the plant food in 5000 pounds of average milk is $4.98.

592. Kind of Litter Used. Manure consists of the excrements of animals mixed with the litter or bedding material that is used to absorb the liquids. These materials vary in the amount of plant food that they contain. The fertilizing value of a ton of wheat straw is $2.40, and of an equal quantity of sawdust it is only $1.60.

593. Proportion of Plant Food Recovered in Manure. Taking into consideration the different kinds of live stock maintained on the average farm and the proportion of growing and mature animals, it may be assumed that three fourths of the nitrogen and phosphoric acid of the feeds and over nine tenths of the potash are recovered in the manure. In a general way it may be said that the manure contains eighty per cent of the fertilizing value of the ration fed and the full value of the materials used for bedding. According to these figures the total value of the manure for a year from a herd of fifty dairy cows fed a daily ration of 10 pounds of grain (corn meal, ground oats, and bran), 35 pounds of corn silage, and 15 pounds of clover hay, and bedded with wheat straw amounts to $2,094.20.

FIG. 232. — Manure pile on the island of Jersey with cistern below to collect liquid manure.

594. Losses in Manures. The foregoing statements refer to fresh manure which has suffered no loss of its valuable constituents. On the average farm, unfortunately, because of lack of care in preventing the losses to which manure is subject, not more than half this value is realized. It will be well to consider these losses and the means by which they may be prevented. The principal ways in which plant food is lost from the manure are as follows: (1) by neglect-ing to save the liquid manure; (2) by loss of ammonia in the stable; (3) by leaching in the barnyard; (4) by hot fermentation.

595. Value of the Liquid Manure. The liquid part of the manure, which is commonly allowed to run away, contains two thirds of the nitrogen and four fifths of the potash excreted by the animal. The following table shows the distribution of the plant food in the manure from the fifty dairy cows mentioned in Section 593.

Value of the solid part	$716.88
Value of the liquid part	1,200.20
Value of the bedding	177.12
Total value of the manure	$2,094.20

It will be seen that if the liquid is lost, the value of the manure will be less than $900 instead of $2,094.20 as calculated. It is evident, therefore, that the stable floors should be made of cement or other watertight material, and that sufficient bedding should be used to absorb all the liquid. In many parts of Europe cisterns are built in connection with the stables to collect the liquid manure; but under American conditions it is best to keep the liquid and solid manures together by the plentiful use of bedding.


FIG. 233. — **Outfit for the distribution of liquid manure on the island of Jersey.**

596. Loss of Ammonia in the Stable. Manure contains enormous numbers of decay bacteria that cause its rapid decomposition. One class of these bacteria liberates ammonia from the liquid manure. Most of the nitrogen in the liquid excrement is in the form of an organic compound called *urea* (CON_2H_4). The bacteria cause the urea to take on water and change to ammonium carbonate, thus:

$$CON_2H_4 + 2H_2O \rightarrow (NH_4)_2CO_3.$$

The ammonium carbonate dissociates and gives off ammonia and carbon dioxide (164), resulting in loss of nitrogen from the manure. The odor of ammonia frequently noticed in the stable is due to this chemical change.

The decomposition does not take place so readily if the liquid is completely absorbed by the bedding. The use of dried muck soil or peat with the bedding is effective in preventing this change. If muck soil is easily obtained, it will pay to dry a few wagon loads of it for use in the stable.

Chemical preservatives are sometimes used in the stable to prevent loss of ammonia. The best material for this purpose is acid phosphate (178). The acid phosphate unites with the ammonia, forming the double salt, calcium ammonium phosphate, which is not volatile; hence this change prevents the escape of ammonia. The calcium sulphate (gypsum) which is always present in commercial acid phosphate is also supposed to be effective in preventing loss of ammonia from the manure, by changing the ammonium carbonate into ammonium sulphate, which is not volatile and does not so readily decompose:

$$(NH_4)_2CO_3 + CaSO_4 \rightarrow CaCO_3 + (NH_4)_2SO_4.$$

The acid phosphate should be dusted over the manure gutter at the rate of one pound a day for each animal. This use of acid phosphate is to be recommended, because, in order to obtain the best results in the field, some phosphate fertilizer should always be used in connection with stable manure (612).

597. Losses in Manure from Leaching. Next to improper absorption of the liquid, the greatest loss in manure comes from leaching by rains. As ordinarily handled the manure is thrown out each day into the open yard to lie for months subjected to washing by the summer or winter rains. In many cases it is even deposited under the eaves of a large barn, and thus the washing process is made more complete. It is absurd to go to the trouble of absorbing all the liquid excrement by means of bedding, and then allow it to be washed out of the manure. The losses in manure due to leaching by rains in the open yard are much greater than most people imagine. Many experiments have been carried on to determine these losses, and the following table gives the results of four such experiments:

FIG. 234. — Great losses result from leaving manure exposed to the weather.

LOSSES IN MANURE FROM LEACHING

PERIOD Days	NITROGEN Per Cent	PHOSPHORIC ACID P_2O_5 Per Cent	POTASH K_2O Per Cent
131	57.0	62.0	72.0
70	44.0	16.0	28.0
76	39.0	63.0	56.0
50	69.0	59.0	72.0
Average	52.2	50.0	57.0

The table shows that the average loss amounted to more than half the plant food in the manure during rather short periods, the longest time being a little over four months.

On many farms the manure is exposed to the weather for a much longer period of time. These losses vary with the climatic conditions and with the quality of the rations. During heavy rains, especially if they occur in warm weather, the losses are much greater than in dry or cold weather. In the experiments noted above, the rainfall was as great during the 50 days of the last experiment as it was in the case of the 131 days of the first one. The relative decrease in value is larger for manures produced from rations of high nutritive value. In other words, the more valuable the manure, the greater will be the percentage of loss from leaching. It is conservative to say that manure exposed to the weather for six months loses fully half its value.

It is worthy of note also that the plant food that is washed out of the manure is the part that is most available, as it is soluble in water and is in the condition in which it can be immediately used by the plants. The manure that remains, on the other hand, represents the tougher and more slowly decomposed material, hence it contains the least available part of the plant food.

FIG. 235. — Covered manure shed with cemented bottom.

Manure is never so valuable as when perfectly fresh.

Even the best methods of handling and care, if the manure is stored, cannot prevent more or less loss of the valuable constituents. For this reason it is advisable to haul the manure directly from the stable to the field each day, whenever the conditions permit.

There are always times on every farm when it is not possible to haul the manure directly to the field, and some suitable place should be provided for its temporary storage (Fig. 235). The essential requisite of such a storage place is that it shall have a cemented bottom to prevent any loss of the liquid manure. A cover to protect it from the rains is desirable. A small manure shed would pay for itself in a single season on a farm maintaining much live stock.

FIG. 236. — Cattle in a covered barnyard.

Some farmers store the manure in what is called a covered barnyard (Fig. 236), which is usually one large room in the barn in which the cattle are allowed to run during the greater part of the day. The floor is cemented and bedding material is liberally used. The cattle tramp the manure into a solid mass and it is allowed to accumulate until it is convenient to remove it to the field. Protected in this way, the manure suffers very little loss of fertilizing constituents.

598. Open Yard Feeding a Wasteful Practice. It is probably true that upon a majority of the farms in America cattle are fed during the winter in open lots, the manure not being hauled away until the following summer or fall, if indeed it is removed at all. This method of feeding presents conditions that result in excessive losses from leaching, and it is safe to say that more than half the fertilizing value of the manure is lost where this practice is followed. In the corn belt of this country, for instance, large numbers of cattle are fed during the winter, and it is not unusual to see a large feeding lot covered to a considerable depth with manure which is spread out and exposed to the weather in such a way that the maximum effects of leaching must take place. There is no doubt that, considered from the fertility point of view alone, these farms would be better off if the corn were sold from the farm and the stover plowed under. The feeding of the future must be done under cover if the fertility of the soil is to be economically maintained.

FIG. 237. — Open yard feeding greatly reduces the fertilizing value of manure.

599. Hot Fermentation. Manure that has been thrown into a loose heap, especially if it contains much horse or sheep manure, soon becomes very hot. The heating sometimes proceeds so far that part of the manure becomes white, or fire-fanged as it is popularly called. An examination will show that ammonia is being evolved in large quantities from the heating manure. This fermentation is caused by certain bacteria that bring about oxidation of the organic matter, the nitrogen being converted into ammonia. The loss of nitrogen in this way is very large, the amount varying in different experiments from 20 per cent to over 80 per cent. In the case of the white fire-fanged material all the nitrogen is driven off. Since the bacteria that cause this rapid oxidation of the organic matter cannot exist in the absence of free oxygen, no heating will take place if the manure pile is so compact that no air can enter it. In a moist, compact manure pile a cold fermentation takes place, which is caused by an entirely different class of bacteria and which does not result in the formation and loss of ammonia. It is evident, therefore, that the stored manure should be carefully compacted as well as protected from leaching. It will be seen that the covered barnyard presents almost ideal conditions for storing manure. A method of storage very similar to the covered barnyard plan has been in use in Europe for many years and is known as the deep stall method. The manure accumulates in the stall in such a way that it is thoroughly packed by the feet of the cattle, and is said to lose very little of its fertilizing value.

600. Composting Manures. Any method of storing manure requires considerable labor, and for that reason storing it is to be avoided in general farming whenever it is possible to use it in the fresh condition. In market gardening, on the other hand, such quantities of manure are used that it is necessary to have it thoroughly rotted before applying, as otherwise the crop would suffer from the heating effect that the large amount of raw manure would have on the soil. While the manure may be rotted by keeping it in a moist, compact heap, it must be remembered that the manure commonly used by market gardeners is the horse manure from the city stables. This heats so rapidly that special care is necessary to prevent hot fermentation, and the pile must be moistened frequently.

Many market gardeners prefer to compost the manure with earth, peat, or muck. This is done by making a foundation of about six inches of dirt and placing on top of this alternate layers of manure and soil, moistening the mass as the heap grows. The sides and the top should be smoothed off and the mass covered with a thin layer of earth to prevent loss of nitrogen. After about two months the pile should be turned over, the materials thoroughly mixed, and more water added, if necessary, to keep the compost moist. A compost in great favor with greenhouse men is one made of manure and sod, these materials being piled in alternate layers as described above. This gives the fibrous compost so desirable for bench and pot work.

Any of the refuse organic materials of the farm or garden may be used in composts. Weeds, refuse parts of plants, dead animals, and kitchen wastes may be added to the manure-earth mixture, or composted separately; for handled in this way they decompose rapidly and without offensive odors. The presence of the earth decreases the loss of ammonia where highly nitrogenous materials are used. In using composts a good practice is to add bone meal to the heap. In this way the plant food in the bone meal is made available to the plants, and the compost is made more valuable.

601. Applying Manure. Two general methods for the application of manure are in common use: one is to throw it into heaps, where it is allowed to remain some time before being spread; the other is to broadcast it directly from the wagon. The first method is objectionable for several reasons. In the first place it increases the work necessary to spread the manure, since it must be handled twice, and it takes no more labor to spread it from the wagon than from the heap on the ground. The leachings from these heaps make the spots directly beneath more fertile than the rest of the field, and hence produce a rank growth at those places (Fig. 238). This uneven growth is undesirable, because in the case of grains it increases the danger of lodging in the more fertile spots; and in any case it results in unevenness in the maturity of the crop. A crop that has a large supply of plant food, for instance, has a longer period of growth than one with a meager supply and consequently is later in maturing. If, therefore, the field is very uneven in fertility, a part of the crop will be ready to harvest some time before the rest has matured. On the other hand, if the manure is spread directly from the wagon, not only is the labor lessened, but the danger of unevenness in growth is to some extent avoided. Moreover there is no likelihood of loss in the value of the manure when it is spread in a thin layer on the ground.

FIG. 238. — **Showing the uneven growth due to allowing manure to remain in heaps before spreading.**

Manure spreaders (Fig. 239) are coming into general use. Some recent experiments seem to indicate that manure gives better returns when spread by the machine than it does when applied by hand. Whatever method is used to spread the manure, it will readily be seen that the finer the material the easier it will be to distribute it evenly.


FIG. 239. — The best way to apply manure is by means of the manure spreader.

602. Where to Use Manure. There is some difference of opinion as to which of the ordinary farm crops give the best returns for the use of stable manure. Probably more farmers use it on the land plowed for corn than in any other way. Corn is especially adapted to utilize the plant food of manure, since it makes the greater part of its growth in midsummer, when nitrification is at its height and the nitrogen of the manure is being made available most rapidly. It is always safe to manure corn heavily.

Many farmers prefer to use manure as a top dressing for grass lands, since such use increases the organic matter of the soil by stimulating the growth of the fibrous roots of the sod. This method gives good returns for the manure used. The permanent pastures should not be neglected, but should be occasionally top dressed with manure and commercial fertilizers.

603. Amount to Apply. Market gardeners use very large quantities of manure, sometimes as much as forty tons to the acre, but they probably use the manure for its physical effect upon the soil even more than for the plant food that it contains. This is partly due to the fact that the gardener cannot conveniently make extended use of green manures as a source of organic matter. For ordinary farm crops, on the other hand, it is not customary to use more than eight to ten tons to the acre, and on general principles it may be stated that somewhat frequent light dressings pay better than very large ones given at long intervals. On the other hand, the amount of manure produced on the average farm is so small when compared with the land to be fertilized that it would be impossible to spread it over all the farm yearly. For this reason it is a good plan to apply the manure to one or two crops in the rotation, thus covering only a part of the farm each year.

604. How Manure Improves Soils. Stable manure adds all the elements of plant food to the soil, and while some of it is not in forms immediately available to the plant it becomes so during a period of years. Manure adds to the soil enormous numbers of bacteria that attack not only the manure itself but the organic matter already in the soil. Manure also improves the physical condition of the soil, and during its decay the acid products formed act upon the potential plant food and make some of it available. Attention has already been called to the fact that when properly handled the manure returns to the soil nearly one half the organic matter of the feeds. Everything considered, manure is probably the best fertilizer the farmer can use, especially when it is reënforced with acid phosphate.

605. Results with Manure. The good results in crop yields from the use of stable manure are known to every farmer and gardener. In one experiment in England the use of manure has maintained the yield of wheat at 34 bushels to the acre for seventy-five years, while the average for the unmanured field was only 13 bushels. At the Ohio

Experiment Station stable manure has given a profit in increased yields of $3.22 for each ton of manure used over a period of twenty years.

FIG. 240. — **The field on the left received ten tons of manure while the one on the right was unmanured.**

Manure differs from other fertilizers in its lasting effects when applied to the soil. In one experiment manure was used on a plot for twenty years, after which its use was discontinued. The good effect was noticeable for more than thirty years after the last application. Every farmer knows that the effects of a single application of manure are evident five or six years after its use.

FIG. 241. — **Using city sewage as a fertilizer in Japan.**

606. City Sewage. Large quantities of plant food are lost in the sewage of the cities. City sewage contains nitrogen, phosphorus, and potassium to the annual value of one dollar for each inhabitant. In China and Japan, two countries noted for their high crop production, the sewage of the cities has been used as a fertilizer for thousands of years. The sewage is carried out to the farms and gardens, diluted with water, and used on the growing crops. Such a procedure is repulsive to the occidental mind;

but it would be a great thing for agriculture if some unobjectionable method could be discovered to make use of the large amount of plant food now being lost in the city sewage.

EXERCISES

Ex. **364**. Is the use of stable manure as a fertilizer of recent or ancient origin? What is the value of the manure produced annually in the United States? Is much of this value lost? Why is a well-kept manure heap an indication of thrift? How is the value of manure stated? What is meant in trade by phosphoric acid? By potash?

Ex. **365**. How do the values of the manure from different kinds of animals compare? Which is the most valuable per ton? To what is this difference in value per ton largely due? Is there much difference in the total value of the manure produced by the different animals from the same feeds?

Ex. **366**. What are the five factors affecting the value of fresh manures? Explain how the value of the manure is affected by the kind of feeds; by the age of the animal; by the kind of animal; by the animal products; by the kind of litter. What proportion of the plant food in the ration is recovered in the manure?

Ex. **367**. If fifty animals were fed the ration given on page 336, what would be the value of the manure produced in a year, assuming that ten pounds of wheat straw were used daily as bedding for each animal? (For the fertilizing constituents of the feeds and bedding see tables in Vivian's First Principles of Soil Fertility.)

Ex. **368**. In what four principal ways is plant food lost from the manure? What proportion of the different elements of plant food is found in the liquid manure? In the solid manure? What proportion of the value of the manure is lost if the liquid is not saved?

Ex. **369**. In what form is the nitrogen in the liquid excrement? What compound do the bacteria form from the urea? Write the reaction. To what is the odor of ammonia in stables due? How may the loss of ammonia in the stable be prevented? Write the equation for the reaction between gypsum and ammonium carbonate. How should acid phosphate be used to prevent loss of ammonia?

Ex. **370**. How much of the plant food is lost when manure is exposed to the weather? How do climatic conditions affect the amount of loss? Explain the statement that the most available part of the plant food is lost by leaching. Why should manure be hauled directly to the field when possible? What are the essentials of a storage place for manure? What is meant by a covered barnyard? Is there much loss in manure stored in a covered barnyard? Are there any farms in your neighborhood where manure is being wasted? Do any of the farms near the school have manure sheds or covered barnyards?

Ex. **371.** Explain why open yard feeding is wasteful of manure. What is meant by hot fermentation of manure? What losses occur from hot fermentation? How may hot fermentation be prevented? What is meant by composting manure? How is a compost heap prepared? What other materials beside manure may be composted?

Ex. **372.** Why should manure never be placed in heaps on the field before spreading? Why is a manure spreader desirable on a farm? Upon what crop is manure most commonly used? What crops do the farmers of your locality manure? Why does corn respond so well to treatment with stable manure? What are the advantages of using manure on grass lands?

Ex. **373.** How much manure is used to the acre by market gardeners? In ordinary farming? Which is the more desirable, heavy applications at long intervals or frequent light applications? Explain how manure improves the soil. Discuss the lasting effect of manure. How much plant food is lost in city sewage? In what countries is the city sewage all saved and used on the crops?

CHAPTER LIX

COMMERCIAL SOURCES OF PLANT FOOD

PHOSPHORUS

607. Phosphorus must be Purchased. Phosphorus is the one element that must be purchased on practically every farm if its fertility is to be maintained. It is the element that is the limiting factor on nearly every farm. Phosphorus is present in small quantities, some of the soils containing as little as 0.01 per cent and very few having as much as 0.15 per cent. From two thirds to three fourths of the phosphorus taken from the soil by the plants is stored in the seeds, and is, therefore, removed from the farm when grain is sold. Animals use phosphorus in making bone and milk; hence the sale of milk or live stock also removes phosphorus from the farm. While the loss of phosphorus from the sale of animal products is much less than that from the sale of the crops, it is sufficiently large to be a decided drain on the phosphorus supply of the soil, especially if the manure produced by the animals does not receive better care than is given to it on the average farm. There is no natural method of increasing the phosphorus of the soil, such as there is in the case of nitrogen, by the fixation due to the nodule-forming bacteria. The farmer, therefore, must purchase phosphorus to replace that which is removed from the farm.

When feeds are purchased in large quantities, as they are on some dairy farms, phosphorus is brought on to the farm in them and may be put into the soil through the manure. In this case, however, much more nitrogen than phosphorus is purchased, and additional phosphorus is necessary if maximum yields are to be obtained. Recent investigations make it clear that a certain balance between the elements of plant food in the soil is essential to the best results in plant growth. In other words, a balanced ration for plants is quite as desirable as a balanced ration for animals.

608. Bone Phosphates. The commercial sources of phosphorus for fertilizers are the bones of animals and the various deposits of mineral phosphates; and, as has been mentioned, the phosphorus is stated in fertilizer trade not as the element, but as phosphoric acid (P_2O_5). The bones of animals have been used as fertilizers for several centuries, and many farmers still prefer them to any other form of phosphorus fertilizers. The mineral matter of bone consists almost entirely of tricalcium phosphate, which is thoroughly permeated by the organic matter of the bone. The bones used in making fertilizers come from the packing houses, and from the reducing establishments which use the animals that die from accident or disease.

Before being used as a fertilizer the bones are ground to a fine powder. If they are ground in the natural condition, the powder is known as *raw bone meal;* if they are ground after they have been steamed to remove the fat, the product is *steamed bone meal.* Raw bone meal contains about 22 per cent of phosphoric acid (9.5 per cent phosphorus) and 4 per cent of nitrogen, while steamed bone meal contains about 28 per cent of phosphoric acid (12 per cent phosphorus) and 2 per cent of nitrogen (179).

The steamed bone meal is the better product to use, since it contains more phosphorus and is more readily decomposed in the soil. The fat in the raw bones interferes with the decay of the bone and has itself no fertilizing value, but can be used to advantage in other ways. Steamed bone meal, too, is usually lower in price than the raw bone meal.

609. Mineral Phosphates. Deposits of tricalcium phosphate are found in several places in this country and in Canada. Most of that used in fertilizers at the present time comes from Tennessee, South Carolina, and Florida. The mineral phosphate has the same chemical composition as that found in the mineral matter of bones, but since it is not permeated with organic matter, as is the bone phosphate, it is less soluble in the soil moisture than is the other.

There is some difference of opinion as to whether plants can utilize the phosphorus of the rock phosphates even when the material is finely ground. It is generally agreed that these phosphates are of practically no value when used on soils that are very low in organic matter; but it is held by many that when mixed with manure or turned under with clover or other green manure crops they are valuable sources of phosphorus. The theory has been advanced that the carbon dioxide evolved by the decaying organic matter makes enough of the phosphate soluble to supply the needs of the growing crop. The advocates of rock phosphate recommend that, as the material is comparatively cheap, it be applied in large quantities (1000 pounds or more every four years) with a green manure crop, and that the soil be kept well supplied with organic matter.

The finely ground rock phosphate is known in some sections of the country as *floats*. A good sample contains at least 28 per cent of phosphoric acid, but as the rock phosphates vary greatly in purity, floats should be purchased only on a guaranteed analysis.

610. Acid Phosphate. It has long been the custom to treat the rock phosphates with sulphuric acid to make the phosphorus more available. The proportions of acid and phosphate used should be such as to convert the tricalcium phosphate into monocalcium phosphate (176 and 178), thus:

$$Ca_3(PO_4)_2 + 2\ H_2SO_4 \rightarrow CaH_4(PO_4)_2 + 2\ CaSO_4.$$

The monocalcium phosphate is soluble in water and hence is available for plant growth. The soluble phosphate is not separated from the calcium sulphate, but the whole mixture resulting from the treatment of the rock phosphate with the acid is sold under the different names of *acid phosphate, superphosphate,* and *acidulated rock.*

If not enough acid is used to convert all the phosphate into the monocalcium form, the following reaction may take place:

$$Ca_3(PO_4)_2 + CaH_4(PO_4)_2 \rightarrow 2\ Ca_2H_2(PO_4)_2.$$

This new compound is dicalcium phosphate, and as it is regarded as an intermediate step in the reversion, or changing back, of monocalcium to tricalcium phosphate, it is known in the trade as *reverted phosphate.* Dicalcium phosphate is not soluble in water, but is readily dissolved by very weak acids and is supposed to be as available to the plants as the monocalcium phosphate. For that reason the phosphoric acid of these two compounds, monocalcium and tricalcium phosphates, is termed *available phosphoric acid.*

Acid phosphates as found on the market contain from 12 to 18 per cent of available phosphoric acid (P_2O_5) in the two forms of monocalcium phosphate and dicalcium phosphate. In all probability more phosphorus is purchased by farmers in the form of acid phosphate than in all other materials combined.

611. Basic slag is used in large quantities in Europe and to some extent in this country as a source of phosphorus for plant food. It is made from certain European iron ores that contain considerable quantities of phosphorus (236). Basic slag contains about 18 per cent of phosphoric acid in a form that is readily available to the crops.

612. Phosphates with Manure. A ton of average stable manure contains about 9 pounds of nitrogen, 2 pounds of phosphorus, and 8 pounds of potassium. Manure, therefore, is evidently deficient in phosphorus, and it is not surprising that the addition of a phosphate to it materially increases its crop-producing power.

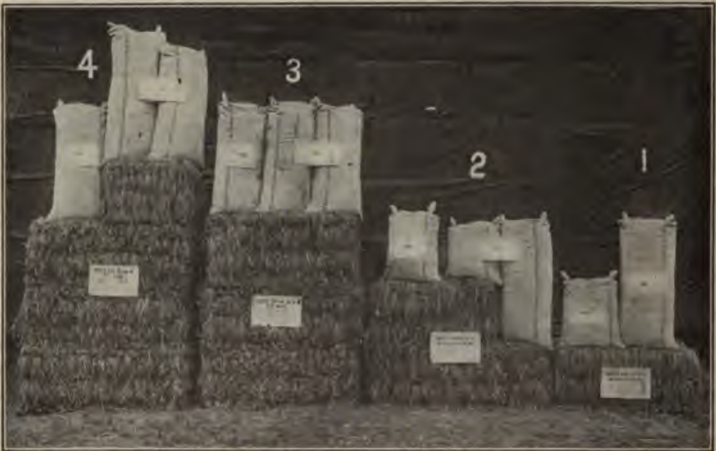

FIG. 242. — Showing the increase in yield from one ton of manure. 1. Yard manure. 2. Stable manure. 3. Stable manure with forty pounds of floats. 4. Stable manure with forty pounds of acid phosphate.

At the Ohio Station an experiment has been running for twenty years in which stable manure alone has been compared with manures to which acid phosphate and floats have been added. All manures were used at the rate of eight tons to the acre on the corn in a three-year rotation of corn, wheat, and clover. In one case forty pounds of acid phosphate and in another forty pounds of floats were added to each ton of manure. As an average for the entire period the stable manure alone gave an increase of crops worth $3.22 for each ton of manure. A ton of stable manure and forty pounds of floats gave a net profit of $4.56, and a ton of manure and forty pounds of acid phosphate gave a net profit of $4.80. It will probably always pay to add some form of phosphate to the manure, and the best way to use it is to scatter it over the manure in the stable. Floats do not have the power of fixing ammonia that was noted in the case of acid phosphate (596).

POTASSIUM

613. Need of Potassium. Most soils are more abundantly supplied with potassium than with phosphorus, and for that reason potassium is less frequently the limiting factor of plant growth. Fully three fourths of the potassium of the mature crop is found in the stems and the leaves, which are not so generally sold from the farm as are the seeds. Likewise the animal retains very little of the potassium of its food; hence most of the potassium of the ration is recovered in the manure.

In the case of peat soils and some sandy soils, however, potassium is the limiting factor and it must be supplied to make them productive. Some plants, such as tobacco, potatoes, and cabbage, require large quantities of potassium, and their successful culture for long periods on most soils necessitates the use of some commercial form of potassium.

614. Potassium Salts. For many years practically all the potash used in fertilizing came from the European potash mines. These mines contain immense deposits of salts, containing various percentages of potash. Only three or four of these products have been commonly used in this country and they are the only ones that will be discussed here.

615. Kainit. This is one of the crude salts which has been ground to a powder. It looks somewhat like common salt, but is darker in color and contains about 12.5 per cent of potash (K_2O) in the form of sulphate, mixed with sulphate and chloride of magnesium.

616. Muriate of potash is manufactured from the crude minerals of the mines by concentration, and contains about 50 per cent of potash in the form of potassium chloride. At the present price the muriate supplies potash at a cheaper price than any of the other materials.

617. Sulphate of potash is another concentrated product of the European mines. What is known as high-grade sulphate contains about 53 per cent of potash in the form of the sulphate (K_2SO_4). The actual potash in this compound costs a trifle more per pound than in the muriate. A lower grade sulphate containing 26 per cent of potash mixed with magnesium sulphate is sold under the name of *double manure salt.*

A relatively small quantity of potash is produced from the ash of the giant kelps which are so abundant on the Pacific coast. There are also certain alkali lakes in this country from which some potash is obtained, notably those in Nebraska. The potash (K_2O) produced from these lakes in 1917 was about 20,000 tons, while the amount annually used in this country previous to 1914 was nearly 300,000 tons.

618. Wood ashes at one time were the sole source of potash for fertilizing purposes, but at present ashes supply only a very small proportion of this element of plant food. Wood ashes vary greatly in composition, the ash from soft woods containing less potash than that from hard woods, the content of potash ranging from 2 to 8 per cent. Potash as found in wood ashes is in a form that is very soluble in water; so that ashes exposed to the weather may have practically all the potash leached out of them (204). Leached ashes as a rule contain less than 2 per cent of potash. As it is not possible to distinguish between leached and unleached ashes by mere physical examination, it is evident that this material should be purchased only from guaranteed analysis.

Nitrogen

619. Importance of Nitrogen. In some ways nitrogen is the most important element of plant food. The more common farm crops use more nitrogen than they do phosphorus and potassium combined. Nitrogen is the most expensive fertilizing material to buy, as it costs about three times as much a pound as either phosphorus or potassium. It is, unfortunately, more easily lost from the farm than any other element of plant food. Approximately two thirds of the nitrogen of the cereal crops is in the seeds and is removed from the farm when grains are sold. It is also lost from the soil in the drainage water and sometimes by denitrification. The ease with which it is lost from manure has been discussed. It has been said that profitable agriculture depends upon an economical method of conserving and renewing the nitrogen supply of the soil. The power of legumes to fix the nitrogen of the air should be utilized as far as possible, and nitrogen should be purchased only as a last resort. Nitrogen fertilizers are so expensive that the matter of their purchase should receive very careful consideration.

The principal sources of nitrogen for fertilizers are: (1) the by-products of the packing houses and rendering establishments, (2) ammonium sulphate from the gas and coke works, (3) nitrate of soda, and (4) cyanamide.

620. Packing House By-products. The more common of the by-products of the rendering, packing, and canning establishments that handle meat and fish are as follows:

Dried blood	containing from	7 to 10	per cent of nitrogen
Meat meal	containing from	12 to 14	per cent of nitrogen
Hoof meal	containing from	11 to 12	per cent of nitrogen
Tankage	containing from	4 to 9	per cent of nitrogen
Dried fish	containing from	8 to 11	per cent of nitrogen

The first three names are self-explanatory. Tankage consists of the refuse material that cannot be used as human food, and which has been placed in a tank and heated with steam under pressure. The heating extracts the fat, which is used in soap making. Tankage is, therefore, variable in character, and in addition to its nitrogen contains from 3 to 12 per cent of phosphoric acid in the form of bone meal.

Most of the fish fertilizers are made from *menhaden*, a fish that is caught in large numbers along the Atlantic coast. The fish are steamed and pressed to extract the oil and the remaining pomace is dried and ground. This material contains from 8 to 11 per cent of nitrogen and about 3 per cent of phosphoric acid. Some of the fish fertilizers consist of the residue of the canning factories, but these are not so valuable as those derived from menhaden.

621. Ammonium sulphate is a by-product in the manufacture of coal gas, animal charcoal, and coke (93). It is richest in nitrogen of all fertilizing materials, containing from 20 per cent to 23 per cent. It gives excellent results on soils that contain plenty of calcium carbonate.

622. Nitrate of soda or Chile saltpeter is a crystalline substance somewhat resembling coarse salt in appearance and is entirely soluble in water. It comes from large deposits in Chile which supply over one million tons of nitrate a year to be used as a fertilizer. Chile saltpeter contains from 15 per cent to 16 per cent of nitrogen in a form that is immediately available to the plant, and for this reason it is the most desirable nitrogenous fertilizer to use when immediate results are desired. It is not held by the soil and consequently should be supplied only as it can be used by the crop.

Calcium nitrate as made at the present time in Norway (148) promises to be an important fertilizing material of the future.

623. Calcium cyanamide ($CaCN_2$) is a new fertilizing material. It is produced by heating calcium carbide (169) to a high temperature in a current of nitrogen until the following reaction takes place:

$$CaC_2 + 2\,N \rightarrow CaCN_2 + C.$$

Calcium cyanamide, called also nitro-lime and lime nitrogen, is a hard gray-black substance resembling coke in appearance and containing about 20 per cent of nitrogen. It decomposes in the soil as follows:

$$CaCN_2 + 3\,H_2O \rightarrow 2\,NH_3 + CaCO_3.$$

Because calcium cyanamide has an injurious effect upon the germination of seeds, it should be applied to the soil some time in advance of seeding, so as to permit the change to ammonia to be completed before the seeds germinate.

624. Low-grade Nitrogen Fertilizers. The high price of nitrogen materials has led to the substitution in some cases of inferior materials for the nitrogen fertilizers described previously. Leather and horn meal, hair and wool wastes, shoddy, dried peat and muck, and garbage tankage all contain nitrogen, but in a condition in which it resists nitrification and is made available very slowly.

625. Relative Availability of Nitrogenous Fertilizers. The percentage of nitrogen present in the different fertilizing materials does not properly indicate their relative fertilizing value. Mention has been made repeatedly of the fact that the plant can make use of the nitrogen only when it is present in the form of nitrates. Nitrate of soda is the only fertilizer on the list that contains nitrogen in the nitrate condition, and consequently is the only one that adds nitrogen to the soil in a form which, without further change, is available to the plant. All the other materials must have their nitrogen converted into nitrates before it can be used by the crop. It must be apparent, therefore, that the value of a nitrogenous fertilizer depends upon both its content of nitrogen and the ease with which it is nitrified.

Of the list mentioned, sulphate of ammonia is the most easily converted into nitrates, provided the soil is abundantly supplied with lime. Next in order comes dried blood.

The nitrogen in dried fish, tankage, hoof meal, and bone meal is readily changed by nitrification and ranks next to blood meal. Horn meal, on the other hand, decomposes very slowly, and the nitrification of leather is so slow as to make it practically worthless as a fertilizer.

Experiments up to date indicate that if nitrate of soda is rated at 100 per cent, the availability of the other materials is as follows:

Nitrate of soda	100
Blood	70
Fish, hoof meal	65
Bone and tankage	60
Leather and wool waste	2–10

626. Suggestions for Using Nitrogen Fertilizers. Two or three suggestions for the selection of nitrogen fertilizers may be drawn from this discussion. For those crops that begin their growth early in the spring, the best results will follow the use of Chile saltpeter, as the soil is likely to be low in nitrates and the process of nitrification is slow at that time. Crops that have very short periods of growth will respond best to nitrogen in nitrates. On the other hand, corn and the other crops that make their growth after the season is well advanced can use the slower acting fertilizers, as can also those crops that occupy the ground permanently. Some farmers prefer to use a fertilizer containing nitrogen in three forms for the crops that grow during the greater part of the season: a little nitrate of soda for immediate use, sulphate of ammonia to supply the nitrogen a little later, and tankage to carry the plant to maturity, these materials being mixed and applied at one time.

EXERCISES

Ex. **374.** Explain why phosphorus must be used on every farm. In what ways is phosphorus removed from the farm? Why is a phosphorus fertilizer needed even when large quantities of feeds are purchased? Do plants need a balanced ration?

Ex. **375.** What are the commercial sources of phosphorus? What is meant by raw bone meal? By steamed bone meal? Why is steamed bone meal the more valuable?

Ex. **376.** What are the sources of the mineral phosphates? Why are the mineral phosphates less available than the bone phosphates? What is the theory regarding the use of floats with organic matter?

Ex. **377.** What is meant by acid phosphate? By what other names is it known? Write the equation for the action of sulphuric acid on mineral phosphate. What may happen if insufficient sulphuric acid is used? What is meant by reverted phosphate? By available phosphoric acid? How much available phosphoric acid do the acid phosphates contain?

Ex. **378.** What is basic slag? How much phosphoric acid does it contain? Is it in an available form? Why is it desirable to use a phosphate with the stable manure? Discuss the results obtained with phosphated manure at the Ohio Experiment Station. How is the phosphate best added to the manure?

Ex. **379.** Why is potassium less likely to be a limiting factor than phosphorus? What soils are most likely to need potassium? Discuss the European salts as a source of potassium for fertilizers. What can you say of the value of wood ashes? Why should wood ashes be protected from the weather?

Ex. **380.** Why is nitrogen sometimes said to be the most important element of plant food? What are the principal sources of nitrogen for fertilizers? Tell what you can about the packing house by-products. What is the source of ammonium sulphate? What is Chile saltpeter? Calcium cyanamide?

Ex. **381.** Are all nitrogen materials equal in value? What is meant by the relative availability of nitrogenous fertilizers? What is the order of their availability? Give some suggestions for using nitrogen fertilizers.

CHAPTER LX

MIXED FERTILIZERS

627. Complete Fertilizers. By far the larger part of the commercial fertilizers used by farmers in this country are purchased in the form known as complete fertilizers. A complete fertilizer, in the sense in which the word is used, is one that contains nitrogen, phosphoric acid, and potash, in proportions that are supposed to be suited to the requirements of farm practice. Almost all these fertilizers are made by mixing two or more of the basic materials heretofore described, the different ingredients being so combined as to give the desired percentage of nitrogen, phosphoric acid, and potash. In case the basic materials alone yield a product that is richer in the essential ingredients than is desired by the manufacturer, sufficient gypsum, dry earth, peat, or other inert matter is added to bring the percentage of these ingredients down to the desired point. Materials added in this way are known as fillers.

628. Home-mixed Fertilizers. If the farmer wishes to do so, he may buy the separate materials previously described and mix them on the farm, and in that way save part of the expense of the manufacture of a complete fertilizer. There are obvious advantages other than economy in home mixing over the purchase of mixed fertilizers. The usual analysis of a mixed fertilizer gives no clew as to the condition or source of the nitrogen, and it is difficult to determine its availability; while in the homemade mixture the condition of the nitrogen can always be known. Home mixing permits the uniting of the different elements in the proportions that have been found best to meet the requirements of the crop and the soil on which it is to be raised, something that is not easily managed with factory mixed fertilizers. By buying the basic materials separately it is possible to apply the different elements at different times, a point that is sometimes of great advantage in feeding a crop, especially if it is one that needs large quantities of nitrogen. In fact the only advantage that can consistently be claimed for the mixed goods is that they are more generally distributed in the market than the basic materials and can, therefore, be more easily purchased in such amounts and at such times as are convenient.

FIG. 243. — Mixing fertilizers on the farm.

A tight barn floor, a straight-edged shovel, and a wire screen, such as is used for screening ashes, are the only requisites for the home mixing of fertilizers (Fig. 243). The mixer weighs the materials and spreads them upon the floor in layers one upon the other. Then beginning at one end he shovels the whole through the screen, repeating the operation three or four times, or until he attains a fairly uniform mixture.

629. Buying Fertilizers. In order to protect the purchaser, most of the states have passed laws compelling the manufacturer to guarantee the amount of plant food in each brand of fertilizer offered for sale. The enforcement of these laws and the chemical examination of the fertilizers to determine if they agree with the guarantee are intrusted to the experiment stations in some states, while in others they are in the hands of the State Department of Agriculture. The results of the analyses are published in bulletins for free distribution, and these should be generally consulted by farmers using fertilizers.

Fertilizers should be purchased absolutely on the basis of the plant food that they contain, and no attention should be paid to the name of the brand, which usually bears no relation to the usefulness of the fertilizer. For the sake of simplicity the analysis of a fertilizer is stated briefly in the following manner: as a 3–10–4, the figures meaning that the fertilizer contains 3 per cent of nitrogen, 10 per cent of available phosphoric acid, and 4 per cent of potash, the ingredients always being stated in the same order.

630. Need of Knowing What Fertilizer Is Required. To buy fertilizers intelligently the farmer must first know the requirements of the crop and the condition of the soil. A soil that is deficient in phosphorus, for instance, will not be much benefited by a fertilizer very low in phosphorus, even though such a fertilizer may be valuable on other land. Unfortunately there is no easy way of determining accurately the immediate fertilizer requirements of a given soil for a particular crop.

631. Need of Studying the Growing Crop. In a general way, the crops themselves may give some valuable suggestions.

(*a*) As a rule, lack of nitrogen is indicated when plants are pale green, or when there is small growth of leaf and stalk, other conditions being favorable.

(*b*) A bright, deep green color with a vigorous growth of leaf or stalk is, in the case of most crops, a sign that nitrogen is not lacking; but such conditions do not necessarily indicate that more nitrogen could not be used to advantage.

(*c*) An excessive growth of leaf or stalk, accompanied by an imperfect bud, flower, and fruit development indicates too much nitrogen for the potash and phosphoric acid present.

(*d*) When such crops as corn, cabbage, grass, and potatoes have a luxuriant, healthful growth, an abundance of potash in the soil is indicated. When fleshy fruits of fine flavor and texture can be successfully grown, the same condition is indicated.

(*e*) When a soil produces good, early maturing crops of grain, with plump and heavy kernels, phosphoric acid will not generally be lacking. A good growth of straw with a small yield of grain, on the other hand, shows that the soil does not contain sufficient phosphorus to balance the nitrogen and potash.

Such general indications may often be helpful, and crops should be studied carefully with these facts in mind.

632. Field Experiments. The only reliable way of ascertaining the proper fertilizer to use on a given field is to compel the soil itself to answer the question. This may be done by an easily conducted field experiment. This experiment consists in dividing a small portion of the field into small plots (Fig. 244) on each of which a different kind of fertilizer is used, the yield being compared with the yield of check plots to which no fertilizing material has been added. Since the different crops vary in their power to extract plant food from the soil, these experiments should cover the entire rotation.

633. Conducting the Experiment. The first important consideration in an experiment of this kind is the selection of the location for the plots. The spot selected should represent as nearly as possible the average condition of the entire field. The soil should be uniform in quality over the entire area devoted to the experiment, so that one may be sure that any difference in yield from the several plots is not due to variation in the composition of the soil. Plots one rod wide and eight rods long will be found a convenient size for this purpose, but those of any size may be used. The simplest experiment that will give any reliable information calls for a row of at least seven plots, with a space of at least 2 feet between each plot. The ground is first plowed and harrowed and then the plots are measured out, each corner being marked by a stake driven well into the ground. The fertilizers for each division are mixed and applied by hand, care being used not to scatter the material beyond the plot for which it is intended. The diagram (Fig. 245) shows the

FIG. 244. — View of a field plot test with fertilizers.

arrangement of the plots and the kind and quantity of fertilizing material to be used on each.

1	No Fertilizer
2	15 lb. Nitrate of Soda 15 lb. Sulphate of Potash 30 lb. Acid Phosphate
3	30 lb. Acid Phosphate 15 lb. Sulphate of Potash
4	No Fertilizer
5	15 lb. Nitrate of Soda 15 lb. Sulphate of Potash
6	15 lb. Nitrate of Soda 30 lb. Acid Phosphate
7	No Fertilizer

FIG. 245. — Arrangement of plots and quantity of fertilizer used in conducting a field experiment.

The plots may be seeded separately, but it saves labor and gives practically as good results if they are planted with the rest of the field. In any case the seeder must be run lengthwise of the plots so as to avoid dragging any of the fertilizer from one plot to another.

634. Harvesting the Crop. The area devoted to the experiment should receive exactly the same treatment during the growing season as the rest of the field, except that in no case is cross cultivation of the plots allowable. If the crop is one that is planted in rows and intertilled, it will be best to harvest the same number of rows from the center of each plot, discarding the outer rows. In the case of small grains, a cord is stretched from stake to stake to outline the plots and the grain is first removed from the intervening spaces. This leaves each plot standing out so distinctly that it can be readily observed and the crop can easily be harvested. The weight of both grain and straw from each plot should be determined.

635. Interpreting the Results. The yield from each of the check plots should be practically the same. If this is the case, it shows that the soil in the area devoted to the experiment is uniform in character. A little thought will show how to decide from the experimental data what elements of fertility give satisfactory results with the crop and soil under investigation. If the yield on all the plots is about the same, for instance, it will be evident that no beneficial results can be expected on that soil from the use of commercial fertilizers. If plot number 2 gives higher results than any of the others, it is to be concluded that nitrogen, phosphoric acid, and potash are all required. If plots 2, 3, and 6 give larger yields than the checks, and 5 does not, the indication is that phosphoric acid alone is necessary. An increased yield on 2, 3, and 5 but not on 6 indicates need of potash. A larger crop on 2, 5, and 6 but not on 3 shows need of nitrogen. A large increase in yield over the checks on 2 and 6 and a smaller increase on 3 and 5 suggest that both nitrogen and phosphoric acid are beneficial but potash is not; and so on.

Experiments similar to these are being conducted by many of the state experiment stations on the different types of soil found in the several states, so as to assist the farmer in determining what kinds of fertilizers are needed for his particular farm.

636. Commercial Fertilizers Not All-sufficient. Absolute dependence should not be placed on commercial fertilizers alone to maintain the fertility of the land. Commercial fertilizers add little or no humus to the soil, and to obtain the best results it is absolutely necessary to provide humus, either by plowing under green crops or by the use of barnyard manure. Numerous experiments have shown that commercial fertilizers give much better returns when used in connection with barnyard manure than if used alone, and they are coming into use in this manner more and more as the subject is more thoroughly investigated.

It may be said in this connection that commercial fertilizers are not merely stimulants, as is frequently imagined, but that they actually supply plant food. If rationally used, they will leave the soil more fertile than before their use, instead of decreasing its fertility, as would happen if a mere stimulant were used. Commercial fertilizers have an important place in the rural economy; but they should not be used to do the work that can better be accomplished by proper husbanding of home resources.

EXERCISES

Ex. **382**. What is meant by a complete fertilizer? How are complete fertilizers made? What is meant by a filler? Find the brand names and the analyses of the fertilizers used by the farmers in your locality.

Ex. **383**. What are the advantages of the home mixing of fertilizers? How should you proceed to make a home-mixed fertilizer? How could you make a 3–10–4 fertilizer from the following materials: Nitrate of soda containing 16 % nitrogen, acid phosphate containing

16 % phosphoric acid, and muriate of potash containing 50 % potash? How much filler would be required?

Ex. **384.** What does a farmer need to know before buying a fertilizer? Upon what basis should the fertilizer be purchased? How much attention should be paid to the name of the fertilizer? Explain how the crop may indicate what fertilizer is needed.

Ex. **385.** Describe a field experiment to determine the fertilizer requirements of a soil. The following results were obtained in an actual experiment similar to the one described: plots 1, 4, and 7 yielded 10 bushels of wheat to the acre. Plots 2, 3, 5, and 6 yielded 26, 19, 13, and 23 bushels in the order named. Were all three elements of fertility needed? Can you tell which element was most needed? Does the experiment show whether there was more need of potash or nitrogen?

Ex. **386.** In another actual experiment plots 1, 4, and 7 produced only 6 bushels of wheat to the acre. Plots 2, 3, 5, and 6 produced 19, 16, 9, and 19 bushels in the order named. What elements are needed in this soil? Which element appears to be most needed?

CHAPTER LXI

TYPES OF FARMING AND FERTILITY

637. Some Fertilizer Always Needed. While soils vary greatly in their original fertility, so that it is customary to speak of rich and poor soils, it is also true that none of them can long produce crops without the addition of some kind of fertilizing material. Any system of farming that is to succeed through a long period of years must provide some method of replacing practically all the plant food removed, no matter how rich the soil was in its virgin condition.

Each farm is in a measure a special study; but the work done by the various agricultural experiment stations makes it possible to formulate a few important generalizations. If the farms are properly drained, if tillage is thorough and rational, if limestone has been added where necessary, and if crops are rotated, the system of fertilization necessary for best results in the various types of farming would be about as outlined in the following sections.

638. All Crops Sold. Where all the crops, including straw and hay, are sold from the farm, the fertility of the soil is quickly exhausted unless fertilizers are freely used. In this case dependence must be placed upon fertilizers that supply nitrogen, phosphorus, and potassium, and a sufficient quantity must be used to replace practically all the plant food removed from the soil. Some difference of opinion exists as to whether the fertilizer should be used wholly on the cereals, or on the hay crop as well. An experiment on this type of farming which has been running for twenty years at the Ohio Experiment Station, an experiment in which the rotation followed was corn, oats, wheat, clover, and timothy and in which the fertilizer was used on the corn, oats, and wheat, gave an annual profit for the use of the fertilizer of $3.53 an acre over and above the cost of the application. The total fertilizing materials used on the three grain crops during each rotation consisted of 480 pounds of nitrate of soda, 320 pounds of acid phosphate, and 260 pounds of muriate of potash. Later experiments indicate that less nitrogen and potassium and more phosphorus would have given greater profit. Under the best of circumstances, however, the total possible profit from this type of farming is comparatively small.

639. Grains Only Sold. In the end it will probably be found more profitable to sell the grain only and return the straw and clover for their manurial effects. A four-year rotation of corn, oats, wheat, and clover will serve as an example. In this type of farming the only loss is in phosphorus, provided the clover, straw, and cornstalks are plowed under; for in this case the fixation of nitrogen by the clover should make up for that sold in the grain. The only commercial fertilizer that need be purchased, therefore, is acid phosphate or some other carrier of phosphoric acid. This type of farming is being recommended for the great grain-growing and cotton-growing sections.

640. Dairy Farming. The type of farming in which fertility is most easily maintained is the one devoted wholly to dairying, in which only such crops are raised as can be fed on the farm. The dairy farmer finds it profitable in milk production to purchase quantities of concentrated feeds, and thereby bring plant food to the farm. To reënforce the manure he should purchase acid [redacted] is preferably used in the stable at the rate day for each animal (612). Under these will increase in fertility if the manure is One run-down farm that was producing corn to the acre was made to yield 85 within eight years by this method of

641. Fat Stock Farming. The same [illegible] apply in this case as in dairy farming; [illegible] feeder usually does not buy as much feed [illegible] dairyman, the farm is not so easily maintain[illegible] of fertility.

642. Mixed Farming. Most of the farm[illegible] are managed by the system known as [illegible] which some of the crops are sold directly [illegible] to the animals and marketed as animal [illegible] this system all the manure should be car[illegible] forced with some phosphate material, a[illegible] land that is to be plowed for corn, prov[illegible] pears in the rotation. Additional pho[illegible] used on the small grains, especially up[illegible] since they respond readily to such treat[illegible] land is not in a high state of fertility it [illegible] profitable to use small quantities of nitro[illegible] on the wheat. At the Ohio Experiment [illegible] tract of forty acres which is being ma[illegible] this system. The rotation is corn, oats, [illegible] with ten tons to the acre of phosphated [illegible] the corn, and chemical fertilizers to the [illegible] years that this plan has been in operation [redacted] several crops have increased as follows: c bushels, oats from 30 to 60 bushels, whe

Pastures. The most neglected part of
ican farm is the area devoted to the per-
. It seems to be assumed that the soil
ture grasses indefinitely without fertiliza-
reas in the older parts of the country show
neglect. The pastures should be as care-
any part of the farm. The best fertilizer
s available is phosphated stable manure
e manure spreader. Permanent pastures
to become acid and, therefore, should
applications of limestone. When stable
ailable chemical fertilizers may be used to
phosphate or bone meal should be used
erate amount of potash salts. The addition
la will encourage the growth of the true
blue grass, but if it is desired to promote
clovers rather than the grasses the nitrate
d. The pastures should be dragged from
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EXERCISES

ly soil produce crops indefinitely without fertiliza-
hings that are fundamental to soil fertility. How
intained if all the crops are sold from the farm?
f the probable profit of this type of farming if long
the fertility be more readily maintained if only
d? What fertilizer would need to be supplied in
. Under what circumstances is this type of farm-

ay the fertility be maintained on dairy farms?
to add a phosphate to the manure even when con-
ased? Are there any dairy or stock farms near

force the manure he should purchase acid phosphate which is preferably used in the stable at the rate of one pound a day for each animal (612). Under these conditions the farm will increase in fertility if the manure is properly preserved. One run-down farm that was producing only 30 bushels of corn to the acre was made to yield 85 bushels to the acre within eight years by this method of farming.

641. Fat Stock Farming. The same general principles apply in this case as in dairy farming; but as the stock feeder usually does not buy as much feeding stuffs as the dairyman, the farm is not so easily maintained in a high state of fertility.

642. Mixed Farming. Most of the farms of this country are managed by the system known as mixed farming, in which some of the crops are sold directly and the rest fed to the animals and marketed as animal products. Under this system all the manure should be carefully saved, reënforced with some phosphate material, and applied to the land that is to be plowed for corn, provided that crop appears in the rotation. Additional phosphorus should be used on the small grains, especially upon wheat if grown, since they respond readily to such treatment. In case the land is not in a high state of fertility it may also be found profitable to use small quantities of nitrogen and potassium on the wheat. At the Ohio Experiment Station there is a tract of forty acres which is being managed according to this system. The rotation is corn, oats, wheat, and clover, with ten tons to the acre of phosphated manure applied to the corn, and chemical fertilizers to the wheat. In the 12 years that this plan has been in operation the yields of the several crops have increased as follows: corn from 34 to 78 bushels, oats from 30 to 60 bushels, wheat from 15 to 34 bushels, and clover hay from 2000 to 6400 pounds to the acre, with indications that the maximum yields have not yet been reached.

A great many of the farms devoted to mixed farming do not produce sufficient manure to make possible the application of ten tons to each acre once in four years. In such cases it will probably be necessary to resort to the occasional use of some form of green manuring if the yields are to be maintained at a high level.

643. Special Crops. There are certain so-called special crops on which it has long been customary to use large quantities of commercial fertilizers. These crops, among which are tobacco, sugar beets, potatoes, onions, cotton, and celery, bring relatively high prices per acre, and consequently are more likely to give profitable returns for heavy fertilization than are the ordinary farm crops. It is quite a common practice to use from 1000 to 2000 pounds to the acre of a high-grade chemical mixture containing all the elements of fertility upon these special crops. While this practice gives good profits under favorable conditions, there is need of much more experimental work with these plants before it can safely be assumed that this is the most economical and profitable scheme of fertilization even for these high-priced crops. Onions and celery, it should be noted, are frequently grown on a muck or peaty soil, in which case the need of potassium in the fertilizer is evident.

644. Market gardening may be said to be one of the most intensive forms of farming. The areas farmed are small, and a large amount of hand labor is involved. The money returns for each acre are relatively large, and the gardener is justified in making large expenditures to maintain fertility. The mechanical condition of the soil is of prime importance, especially in the production of the root crops, such as radishes, carrots, parsnips, and beets. The best practice is to use composted manure in large quantities, supplementing it with green manures, if the supply of stable manure is not sufficient to maintain a high percentage of organic matter in the soil. The Chinese, who are excellent gardeners, grow clovers and other legumes which they compost with earth and use in lieu of composted stable manure. Since it is becoming increasingly difficult for most gardeners to secure stable manure, this practice of the Chinese gardener may be found to be useful in this country under some conditions. Phosphoric acid and potash fertilizers should be used in abundance, preferably in the compost, since they are retained by the soil, and extra nitrogen should be supplied in the form of nitrate of soda as needed by the crops. Nitrate of soda is especially valuable for the leafy crops that make their growth in the early spring, such as spinach, lettuce, early cabbage and cauliflower, asparagus, and rhubarb. The prices obtained from these crops depend largely upon their earliness, and nitrate of soda forces them into rapid growth. Peter Henderson tells of one case in which an acre of very early cauliflower sold for $1000, while an adjoining acre which, because of improper fertilization, was two weeks later in reaching the market brought only $200.

645. Orcharding. In too many cases the orchard receives no fertilizer of any kind, and yet on many soils the orchard fruits give handsome returns for fertilization. Either manure, commercial fertilizers, or both may be used with good results. Cover crops, which can be plowed under or used around the trees as a mulch, are of assistance in fertilizing an orchard. The frontispiece shows how readily apple trees respond to fertilizers on some soils.

646. Permanent Pastures. The most neglected part of the average American farm is the area devoted to the permanent pasture. It seems to be assumed that the soil can produce pasture grasses indefinitely without fertilization, and these areas in the older parts of the country show the effect of such neglect. The pastures should be as carefully fertilized as any part of the farm. The best fertilizer to use when it is available is phosphated stable manure applied with the manure spreader. Permanent pastures have a tendency to become acid and, therefore, should receive occasional applications of limestone. When stable manure is not available chemical fertilizers may be used to advantage. Acid phosphate or bone meal should be used freely with a moderate amount of potash salts. The addition of nitrate of soda will encourage the growth of the true grasses, such as blue grass, but if it is desired to promote the growth of the clovers rather than the grasses the nitrate should be omitted. The pastures should be dragged from time to time to distribute the droppings of the animals, and they are also benefited by an occasional clipping with the mowing machine.

EXERCISES

Ex. 387. Will any soil produce crops indefinitely without fertilization? Name four things that are fundamental to soil fertility. How can fertility be maintained if all the crops are sold from the farm? What can you say of the probable profit of this type of farming if long continued? Could the fertility be more readily maintained if only the grains were sold? What fertilizer would need to be supplied in that case? Explain. Under what circumstances is this type of farming recommended?

Ex. 388. How may the fertility be maintained on dairy farms? Why is it desirable to add a phosphate to the manure even when concentrates are purchased? Are there any dairy or stock farms near the school? Is the manure properly cared for on these farms? Ascertain whether any of the farmers use phosphate with the manure.

Ex. 389. What is meant by mixed farming? If a farm used a rotation of corn, oats, wheat, and clover, and all the wheat and half the corn and oats were sold, and the other materials were used on the farm, what plan of fertilization should you recommend for the farm? Explain why. When is green manuring advisable in mixed farming?

Ex. 390. Tell what you can about fertilizers for special crops. If any of the special crops named in the text are grown in your neighborhood report on the fertilizer used on them. Why is the mechanical condition of the soil of prime importance in market gardening? What is the best way to improve the mechanical condition of the soil? Why is nitrate of soda especially valuable to the market gardener? In case stable manure cannot be procured how may the amount of compost be increased?

Ex. 391. Does it pay to fertilize the orchard? What method may be used? What is meant by a cover crop? Are there any orchards in your vicinity in which cover crops are used? What kinds? What can you say about the need of fertilizing the permanent pastures? What is the best fertilizer for pastures? What chemicals should you recommend if blue grass was especially desired? What other treatment should you suggest for the pastures?

APPENDIX

(a) LIST OF CHEMICALS NEEDED FOR A CLASS OF TWELVE

These chemicals may be of the grade known as "Pure" and need not be of the more expensive C. P. grade.

2 lb. Acid, acetic
4 oz. Acid, arsenious
2 oz. Acid, benzoic
4 oz. Acid, boric
4 oz. Acid, carbolic
4 oz. Acid, citric, crystals
6 lb. Acid, hydrochloric
2 oz. Acid, lactic
6 lb. Acid, nitric
4 oz. Acid, oxalic, crystals
4 oz. Acid, phosphoric
1 oz. Acid, salicylic
18 lb. Acid, sulphuric, commercial
9 lb. Acid, sulphuric, pure
2 oz. Acid, tannic
4 oz. Acid, tartaric, crystals
1 gal. Alcohol, denatured
1 pt. Alcohol, grain
1 qt. Alcohol, wood
8 oz. Alum, ammonium
8 oz. Alum, chrome
8 oz. Alum, ferric
8 oz. Alum, potassium
1 lb. Aluminum sulphate
1 lb. Ammonium carbonate
1 lb. Ammonium chloride
8 lb. Ammonium hydroxide
2 oz. Ammonium molybdate
1 lb. Ammonium nitrate
1 lb. Ammonium sulphate
8 oz. Barium chloride
8 oz. Barium hydroxide
8 oz. Barium nitrate
1 can Bleaching powder
1 lb. Calcium carbide
8 oz. Calcium carbonate, precipitated
1 lb. Calcium chloride
1 lb. Calcium fluoride
1 oz. Calcium metal
1 lb. Calcium phosphate
1 lb. Calcium sulphate
2 lb. Carbon bisulphide
1 lb. Carbon tetrachloride
1 lb. Carborundum
1 lb. Charcoal, animal
1 lb. Charcoal, wood
8 oz. Chloroform
1 oz. Cobalt nitrate
2 oz. Cochineal bugs
4 oz. Collodion
4 oz. Copper oxide, coarse
5 lb. Copper sulphate
8 oz. Copper turnings
4 oz. Cresol
1 lb. Dextrose, lump
1 lb. Ether
5 lb. Formaldehyde
8 oz. Gall nuts
1 lb. Glycerin
4 oz. Gum arabic
4 oz. Gum damar
4 oz. Hydrogen peroxide
2 oz. Iodine
4 oz. Iron ammonium citrate
1 lb. Iron pyrites
1 lb. Iron sulphate
1 lb. Iron sulphide

(a) LIST OF CHEMICALS NEEDED (*Continued*)

1 lb. Kaolin
4 oz. Lampblack
1 lb. Lead acetate
1 lb. Lead arsenate
1 lb. Lead metal
1 lb. Lead oxide, red
1 lb. Lead oxide, yellow
5 vials Litmus paper, blue
5 vials Litmus paper, red
4 oz. Litmus solution
4 oz. Logwood extract
4 oz. Magnesium chloride
2 oz. Magnesium powder
2 oz. Magnesium ribbon
8 oz. Magnesium sulphate
1 oz. Maltose
2 lb. Manganese dioxide
2 lb. Mercury
2 oz. Mercury bichloride
1 lb. Naphthalene
1 lb. Oil, cottonseed
1 lb. Oil, linseed, boiled
1 lb. Oil, linseed, raw
8 oz. Oil, olive
1 oz. Oil, wintergreen
1 oz. Pancreatin
2 lb. Paraffin
8 oz. Paris green
2 oz. Pepsin, scale
1 lb. Petrolatum
1 oz. Phenolphthalein
2 oz. Phosphorus, red
2 oz. Phosphorus, yellow
1 oz. 5% sol. Platinum chloride
1 lb. Potassium bichromate
1 lb. Potassium bitartrate
2 oz. Potassium bromide
1 lb. Potassium carbonate
2 lb. Potassium chlorate
8 oz. Potassium chloride
2 oz. Potassium cyanide
4 oz. Potassium ferricyanide
8 oz. Potassium hydroxide, sticks
2 oz. Potassium iodide
1 oz. Potassium metal
1 lb. Potassium nitrate
8 oz. Potassium permanganate
1 lb. Potassium sulphate
2 lb. Rochelle salts
1 lb. Rosin, common
4 oz. Shellac
2 oz. Silver nitrate
1 lb. Soda lime
1 lb. Sodium acetate
4 oz. Sodium arsenite
2 oz. Sodium benzoate
1 lb. Sodium bicarbonate
1 lb. Sodium borate
1 lb. Sodium carbonate
4 lb. Sodium hydroxide, sticks
2 lb. Sodium hyposulphite
2 oz. Sodium metal
2 lb. Sodium nitrate
4 oz. Sodium nitrite
8 oz. Sodium peroxide
1 lb. Sodium phosphate
1 gal. Sodium silicate
1 lb. Sodium sulphate
1 lb. Sodium sulphite
2 lb. Sulphur, flowers of
8 oz. Sulphur, precipitated
2 lb. Sulphur, roll
1 lb. Turpentine
1 oz. Vermilion
5 lb. Zinc, mossy
8 oz. Zinc chloride
1 lb. Zinc dust
1 lb. Zinc oxide
1 lb. Zinc sulphide

(*b*) SUBSTANCES WHICH MAY BE OBTAINED LOCAI AS REQUIRED

- Absorbent cotton
- "Ammo" solid ammonia
- Ammonia, household
- Candles
- Chloride of lime
- Coal, hard
- Coal, soft
- Coke
- Cream of tartar
- Dry cells
- Eggs
- Fertilizing materials
- Gasoline
- Gelatin, Jello
- Grain products:
  - Barley
  - Corn meal
  - Dextrin
  - Flaxseed
  - Malt
  - Oats
  - Various starches
  - Wheat flour
- Graphite
- Insect powders — pyrethrum
- Insecticides, nicotine
- Iron, cast
- Iron, wrought
- Karo sirup
- Kerosene
- Lime
- Limestone and marble
- Milk products:
  - Butter
  - Buttermilk
  - Condensed milk
  - Powdered milk
  - Renovated butter
  - Skim milk
  - Whole milk
- Miscellaneous fats:
  - Cocoanut oil
  - Cottolene
  - Crisco
  - Lard
  - Oleomargarine
  - Tallow
- Paints and varnishes
- Paraform or formacone
- Petroleum, crude
- Picture frame wire
- Plaster of Paris
- Portland cement
- Raisins
- Rennet extract
- Salt
- Sapolio and scouring soaps
- Soap powders
- Soaps, various
- Steel
- Sugar
- Textiles:
  - Cotton cloth
  - Linen cloth
  - Mixed cotton and wool
  - Silk cloth
  - Woolen cloth
- Vinegars
- Wood ashes
- Yeast, compressed

(*c*) APPARATUS FOR GENERAL USE

Figure numbers refer to illustrations of the apparatus in the text.

Homemade apparatus may be prepared as substitutes for many of the articles in lists *c* and *d*.

Balance — Capacity 2 pounds sensitive to 0.1 gram (Fig. 59)
Apparatus for electrolytic dissociation of water (Fig. 49)
Eudiometer (Fig. 54)
Air pump (Fig. 31)
Microscope
Water oven (Fig. 142)
Babcock testing outfit with glassware for milk, skim milk, and cream (Figs. 164–170)
Quevenne lactometer (Fig. 171)
Truog soil tester (Fig. 226)
Graham-McCall drainage apparatus (Fig. 206)
Soil auger (Fig. 193)
Gasoline blast torch (Fig. 5)
Magnet
Set of cork borers (Fig. 13)
2 matched F. thermometers (Fig. 40)
1 centigrade thermometer (Fig. 30)
Separatory funnel (50 cc.) (Fig. 84)
Small piece platinum wire
Large glass tubes or lamp chimneys (Fig. 204)
One quart sprayer for oat smut exercise
Tall glass cylinder
Small combustion burner (Fig. 44)
Mortar and pestle

(*d*) APPARATUS REQUIRED FOR EACH STUDENT WHO IS TO PERFORM THE EXPERIMENTS

10 ft. small glass tubing, 5 mm.
Bunsen burner or alcohol lamp (Figs. 1, 2)
Corks, common and 2-hole rubber
6 test tubes, common
Test tube, hard glass
Test tube brush
Evaporating dish, 2 in. (Fig. 27)
Florence flask, 500 cc. (Fig. 25)
Florence flask, 1000 cc. (Fig. 12)
Condenser (Fig. 28)
Bottles of various sizes
Watch glass, 2½ in. (Fig. 37)
1 ft. hard glass tubing, 1½ cm.
2 beakers (500 cc. and 250 cc.) (Fig. 15)
Pinchcock
$CaCl_2$ drying tube (Fig. 48)
Thistle tube (Fig. 46)
Funnel, 2½ in. (Fig. 20)
Crucible, porcelain (Fig. 83)
Graduated glass cylinder, 100 cc. (Fig. 172)
Triangular file
Rat-tail file
Blow pipe (Fig. 125)
Ring stand with rings and clamps (Fig. 20)
Wire gauze
Glass plates, 3 in. × 3 in.
Asbestos paper

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MAR 27 1922
APR 4
AUG 10
SEP 26 '2
FEB 19
OV 24 '2
UG 24 '2
Jan 3
April 5
Apr 9

NAME	DATE DUE
Bernice Dimick	MAR 27 1922
S. Burdick	APR 4 1922
B. Dimick	MAY 1 1922
Geo. Stevens per H.O.G.	AUG 10 '22
S.S. White	SEP 26 '23
S.S. White	FEB 19 '24
I. Ladden	NOV 24 '24
Geo. Henshaw Gardiner	AUG 24 '25
Nilona K. Fuelling	Jan 3
Mildred Fuechu	Apr. 5
Mildred Fuechu	Apr 8
A. Van Holand	
A. Welten	
Miss Oph	

INTERNATIONAL ATOMIC WEIGHTS (1914)

O = 16

Aluminum	Al	27.1
Antimony	Sb	120.2
Argon	A	39.88
Arsenic	As	74.96
Barium	Ba	137.37
Beryllium	Be	9.1
Bismuth	Bi	208.0
Boron	B	11.0
Bromine	Br	79.92
Cadmium	Cd	112.40
Cæsium	Cs	132.81
Calcium	Ca	40.07
Carbon	C	12.00
Cerium	Ce	140.25
Chlorine	Cl	35.46
Chromium	Cr	52.0
Cobalt	Co	58.97
Columbium	Cb	93.5
Copper	Cu	63.57
Dysprosium	Dy	162.5
Erbium	Er	167.7
Europium	Eu	152.0
Fluorine	F	19.0
Gadolinium	Gd	157.3
Gallium	Ga	69.9
Germanium	Ge	72.5
Gold	Au	197.2
Helium	He	3.99
Holmium	Ho	163.5
Hydrogen	H	1.008
Indium	In	114.8
Iodine	I	126.92
Iridium	Ir	193.1
Iron	Fe	55.84
Krypton	Kr	82.92
Lanthanum	La	139.0
Lead	Pb	207.10
Lithium	Li	6.94
Lutecium	Lu	174.0
Magnesium	Mg	24.32
Manganese	Mn	54.93
Mercury	Hg	200.6

Molybdenum	Mo	96.0
Neodymium	Nd	144.3
Neon	Ne	20.2
Nickel	Ni	58.68
Niton (radium emanation)	Nt	222.4
Nitrogen	N	14.01
Osmium	Os	190.9
Oxygen	O	16.00
Palladium	Pd	106.7
Phosphorus	P	31.04
Platinum	Pt	195.2
Potassium	K	39.10
Praseodymium	Pr	140.6
Radium	Ra	226.4
Rhodium	Rh	102.9
Rubidium	Rb	85.45
Ruthenium	Ru	101.7
Samarium	Sa	150.4
Scandium	Sc	44.1
Selenium	Se	79.2
Silicon	Si	28.3
Silver	Ag	107.88
Sodium	Na	23.00
Strontium	Sr	87.63
Sulphur	S	32.07
Tantalum	Ta	181.5
Tellurium	Te	127.5
Terbium	Tb	159.2
Thallium	Tl	204.0
Thorium	Th	232.4
Thulium	Tm	168.5
Tin	Sn	119.0
Titanium	Ti	48.1
Tungsten	W	184.0
Uranium	U	238.5
Vanadium	V	51.0
Xenon	Xe	130.2
Ytterbium (Neoytterbium)	Yb	172.0
Yttrium	Yt	89.0
Zinc	Zn	65.37
Zirconium	Zr	90.6

INTERNATIONAL ATOMIC WEIGHTS (1914)

O = 16

Aluminum	Al	27.1	Molybdenum	Mo	96.0
Antimony	Sb	120.2	Neodymium	Nd	144.3
Argon	A	39.88	Neon	Ne	20.2
Arsenic	As	74.96	Nickel	Ni	58.68
Barium	Ba	137.37	Niton (radium emanation)	Nt	222.4
Beryllium	Be	9.1	Nitrogen	N	14.01
Bismuth	Bi	208.0	Osmium	Os	190.9
Boron	B	11.0	Oxygen	O	16.00
Bromine	Br	79.92	Palladium	Pd	106.7
Cadmium	Cd	112.40	Phosphorus	P	31.04
Cæsium	Cs	132.81	Platinum	Pt	195.2
Calcium	Ca	40.07	Potassium	K	39.10
Carbon	C	12.00	Praseodymium	Pr	140.6
Cerium	Ce	140.25	Radium	Ra	226.4
Chlorine	Cl	35.46	Rhodium	Rh	102.9
Chromium	Cr	52.0	Rubidium	Rb	85.45
Cobalt	Co	58.97	Ruthenium	Ru	101.7
Columbium	Cb	93.5	Samarium	Sa	150.4
Copper	Cu	63.57	Scandium	Sc	44.1
Dysprosium	Dy	162.5	Selenium	Se	79.2
Erbium	Er	167.7	Silicon	Si	28.3
Europium	Eu	152.0	Silver	Ag	107.88
Fluorine	F	19.0	Sodium	Na	23.00
Gadolinium	Gd	157.3	Strontium	Sr	87.63
Gallium	Ga	69.9	Sulphur	S	32.07
Germanium	Ge	72.5	Tantalum	Ta	181.5
Gold	Au	197.2	Tellurium	Te	127.5
Helium	He	3.99	Terbium	Tb	159.2
Holmium	Ho	163.5	Thallium	Tl	204.0
Hydrogen	H	1.008	Thorium	Th	232.4
Indium	In	114.8	Thulium	Tm	168.5
Iodine	I	126.92	Tin	Sn	119.0
Iridium	Ir	193.1	Titanium	Ti	48.1
Iron	Fe	55.84	Tungsten	W	184.0
Krypton	Kr	82.92	Uranium	U	238.5
Lanthanum	La	139.0	Vanadium	V	51.0
Lead	Pb	207.10	Xenon	Xe	130.2
Lithium	Li	6.94			
Lutecium	Lu	174.0	Ytterbium (Neoytterbium)	Yb	172.0
Magnesium	Mg	24.32	Yttrium	Yt	89.0
Manganese	Mn	54.93	Zinc	Zn	65.37
Mercury	Hg	200.6	Zirconium	Zr	90.6